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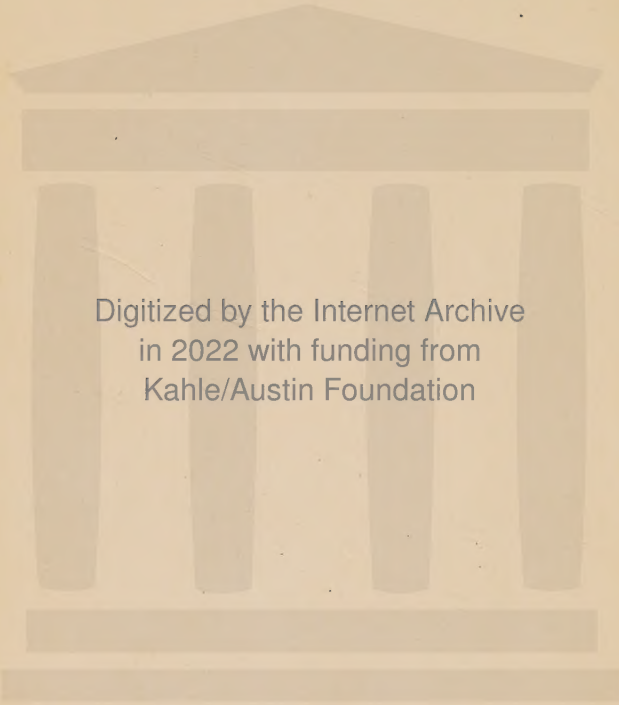
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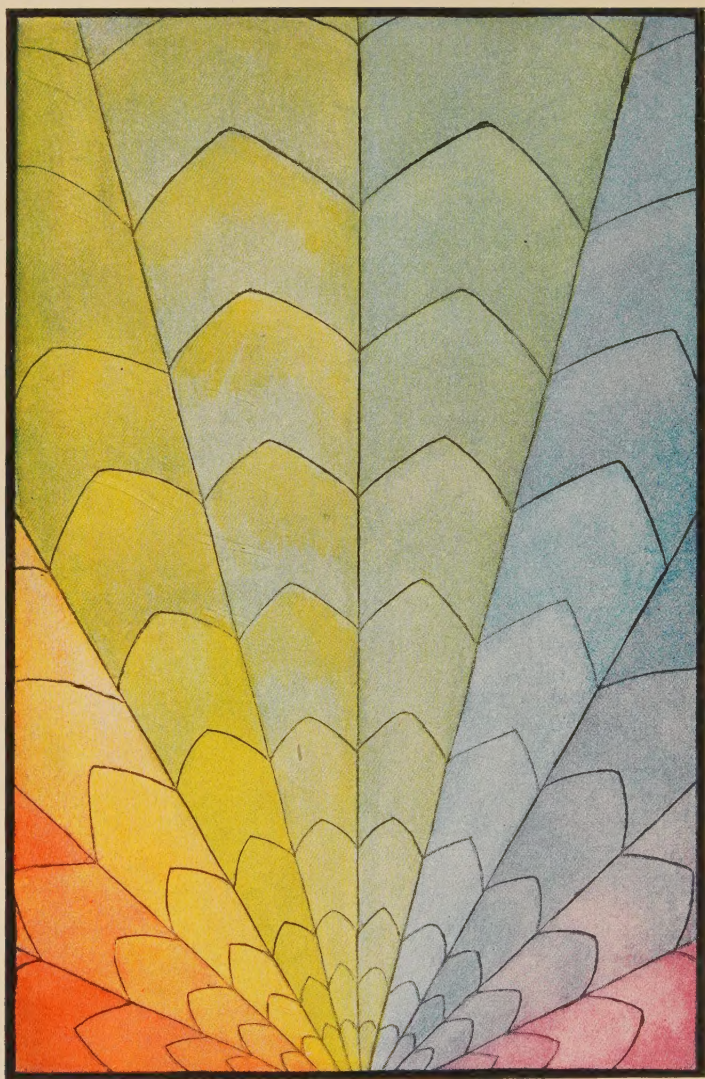
GEORGE EDWARD MAYCOCK

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From coal tar the chemist extracts all the colors of the rainbow
(Frontispiece)

BEGINNING CHEMISTRY

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FLETCHER, SMITH, AND HARROW
BEGINNING CHEMISTRY

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PREFACE

THERE are so many books dealing with elementary chemistry, that there seems little justification for adding another to the list unless some sound reasons can be advanced for such an addition. Twenty years of experience in high school teaching have given the authors certain definite views as to the requirements of pupils; and these views find expression in this book.

In the first place, we are of the opinion that the approach to the subject should be more gradual than is usually the case. Most boys and girls taking up chemistry have already had some elementary science; and it is well, if possible, to make use of such knowledge. The first six chapters have been written with this in mind. These chapters, we feel, are unique in as much as we believe they represent a very gradual transition from what the pupils already know to what they are expected to master.

We proceed, in the next two chapters, to a study of matter from the modern point of view, with the very definite idea in mind that the electron is not merely a chemical curiosity, but a valuable tool to be used throughout the text; and hence to be introduced very early in the course.

The book is an elementary book — an introduction to the subject. The importance of using in an elementary book simple language, language shorn of all ambiguity, and within the range of the pupil, cannot be over-emphasized. This has constantly been kept in mind.

But, elementary though the book is, we have availed ourselves of current chemical literature in order to include important innovations. The chapters on The Colloidal Condition, Chemistry in Agriculture, Foods and Organic Chemistry are particularly rich in

such material. The arrangement of the subject matter, and therefore the titles of chapters, is somewhat different from the conventional type of textbook. For instance, since explosions are common phenomena in chemical change, it has seemed to the authors that it would be better to group together all the topics that come under this head, rather than study gas engine phenomena under gasoline, and explosives under nitrogen compounds. The chapter on Organic Chemistry is particularly comprehensive; and here, as in some other cases, what and how much material is to be used depend upon the individual preferences of the teacher and the capabilities of the pupils.

Our experience has taught us that where summaries at the end of each chapter are used, pupils are inclined to read these summaries and little else, receiving in this way, an incomplete picture of the entire subject. For this reason the questions at the end of each chapter have been devised with the idea not merely of covering the material, which they do, but of stimulating thought. To answer many of these questions it is not always easy for the pupil to trust to his memory or for him merely to turn back a few pages to get the answer. He has to think the thing out.

For review purposes we have added at the end of the book brief summaries; and these, together with the questions, are, in our opinion, more than the equivalent of the usual type of summary.

For the pupils who show exceptional interest and ability in this subject optional questions have been devised. Answers to these questions often require the use of reference books or articles in current literature; and in this connection the references given at the end of each chapter may be used to great advantage.

A word with regard to illustrations. Our experience convinces us that photographs of mechanical applications of chemical principles reproduced, as they are of necessity, on an extremely small scale, give the pupil very little information. The student who has never seen a blast furnace gets little or no adequate impression from the picture of one. He has actually to see one to

gain an impression. But, in any case, since our object in this introductory course in chemistry is not to make technicians of the pupils, but rather to make them familiar with a few of the more important principles and their applications, sectional drawings, emphasizing the *chemistry* rather than the *machinery* involved, have been introduced.

We have covered in this book the syllabus prepared by the American Chemical Society as well as the requirements of the College Entrance Board and other important syllabi. But we have not limited ourselves to these requirements. The book is an introduction to chemistry in its broadest sense, and teachers will find sufficient material to take care of any reasonable requirement.

We believe that a complete and carefully prepared index is of the utmost importance in any text in chemistry and is of considerable value to the pupil. Care has been taken in the preparation of this index to see that it is in harmony with the recognized standards for chemistry books.

Special acknowledgment is given for the use of tables adapted from Sherman's *Chemistry of Food and Nutrition* (1926), The Macmillan Company, and from Stoughton's *The Metallurgy of Iron and Steel*, McGraw-Hill Book Company New York.

It merely remains for us to add that we shall be grateful to our colleagues for any suggestions for improvement in content and method which they may have to offer.

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International Newsreel

FIG. 1. Egypt, the probable birthplace of chemistry.

CHAPTER I

INTRODUCTION

1. The word “chemistry.” It is supposed that the word “chemistry” has something to do with the Greek word “Khemia,” which means “Egypt.” One guess has it that the science was first systematically cultivated by the ancient Egyptian priests who, we know, dabbled in the art of preparing various medicines, and who, therefore, must have studied the properties of various substances (which is what chemistry deals with).

2. What is chemistry? We have just said that chemistry deals with the properties of substances. We may add that it really deals with the properties of everything about us, whether man or stone, including things too big or too small to be seen with the naked eye. Its more particular province is to study the changes which matter undergoes. Some of these changes are discussed in the next paragraph.



FIG. 2. Locomotive.

3. The achievements of chemistry. People are beginning to call our era *the age of chemistry*. Consider some of the achievements due to chemists and you will understand why "the age of chemistry." Out of iron ore and out of coke the chemist has taught us how to make steel, which, in turn, makes railways and steamships and modern buildings possible. Out of petroleum he gets gasoline, the fuel of your automobile. From coal he extracts coal-tar, which has become the starting point for the manufacture of all the colors of the rainbow and most of the drugs on the market. From the air he plucks nitrogen, combines it with several elements, spreads the product over the soil, and the earth becomes fertile as never before, yielding crops in abundance. In his laboratory the chemist builds up perfumes which rival the choicest products of the flower. From wood pulp worth five cents a pound he now makes artificial silk worth \$2.50 a pound. He has almost learned how to dispense with the tree which gives him and the world its rubber. He has extracted and isolated such substances as adrenalin and thyroxin and insulin, with the help of which millions of



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FIG. 3. Steamship made out of steel.

sufferers are put on their feet again. This same chemist has studied the foods we eat to such good purpose that he has discovered the vitamins they contain and the important part which these vitamins play in the life and health of the people of the world.

To what an extent chemistry enters into diverse fields of human endeavor is strikingly illustrated in the flight

undertaken by Lindbergh in crossing the Atlantic. The engines of his airplane were composed largely of high-grade alloy steel, which had first to be made by the chemist; the crankshaft and other parts vital to the mechanism contained substantial per-

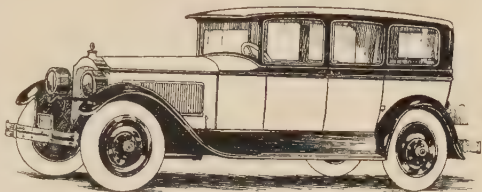


FIG. 4. Automobile made out of steel.

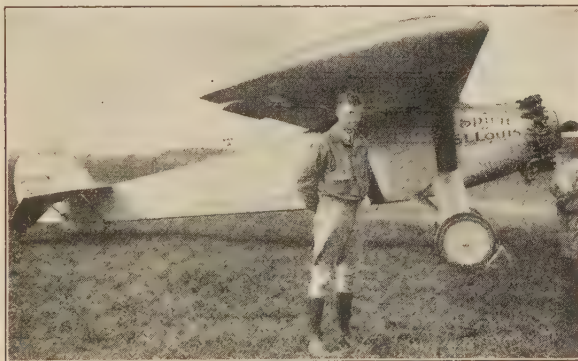


FIG. 5. Lindbergh's plane.

Wide World

centages of chromium, vanadium, and nickel, all products of the chemist's laboratory; and the pyrex glass, which contributed materially to the success of the radio communication, is one of the latest productions due to the chemist.

4. Chemistry in industry. The preceding paragraph must have suggested to you that chemistry plays an important part in indus-

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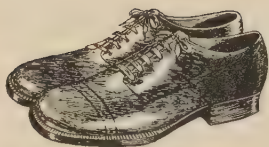


FIG. 6. Shoes made from leather.

try. It has, indeed, been estimated that more than one-half of the manufacturers in this country are dependent upon chemical processes. The chemist is involved in the making of textiles; in the manufacture of iron and steel and gasoline; in the production of leather, paper, ceramics, and glass; in the preparation of countless food products; in the manufacture of dyes and paints and varnishes and soaps and sugar — to name but some of the commercial products.

5. Chemistry the enemy of waste. The ideal in industry — as also in life, for that matter — is to achieve the maximum result with the minimum of effort. Waste, in one form or another, is the enemy that forever attempts to prevent us from achieving our goal. It is not one of the smallest triumphs of chemistry to have shown — and to show — how waste can be avoided. A few examples will illustrate what we have in mind. The life of a locomotive boiler, costing \$1700, was two years.

After the water which it uses had first been properly treated to prevent corrosion, the boiler lasted fourteen years. In this way its life was lengthened sevenfold and the coal consumed by the railroads was reduced by many millions of tons per year. Take another example. Studies by chemists dealing with the lighting problem have produced types of lamps yielding more light for less money. To put this in a more graphic way: if we in this country had used as much

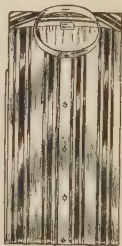


FIG. 7. Shirt made of rayon.



FIG. 8. Glassware.

light in 1912 as we did in 1922, we would have had to pay one and a half billion dollars more for it, and we would have used 25,000,000 tons more coal. And here is a third example: in making yeast, the use of a better grade of mineral salts prevents a

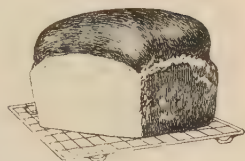


FIG. 9. A loaf of bread.

yearly waste in the fermentation process of 200,000 barrels of flour, 30,000 barrels of sugar, and 15,000,000 pounds of yeast. These savings reach the total of \$5,000,000 a year.



FIG. 10. Mazda lamp.

6. Chemistry and chemists. The avoidance of waste in industry, which has just been discussed, and the stupendous achievements in chemistry, which have also been alluded to, have been made possible by the careful search to discover facts and by a critical investigation of the problems before us

(which is what we mean by "research" in science). Critical to the point of never having an opinion unless it is backed by facts taken from his laboratory, unassuming, seldom heard from and seldom wishing to be heard, with but one passion that absorbs him,



FIG. 11. Part of a boiler badly corroded.

and that is to get as *near the truth* of the matter as his studies will permit, the chemist works in his laboratory, sending forth a "fact" here and a "fact" there, which, all together, have materially contributed towards establishing the civilization of our time.



FIG. 12.

This should teach us a lesson, and, incidentally, should give us a most valuable reason for studying chemistry. In a large measure chemistry has

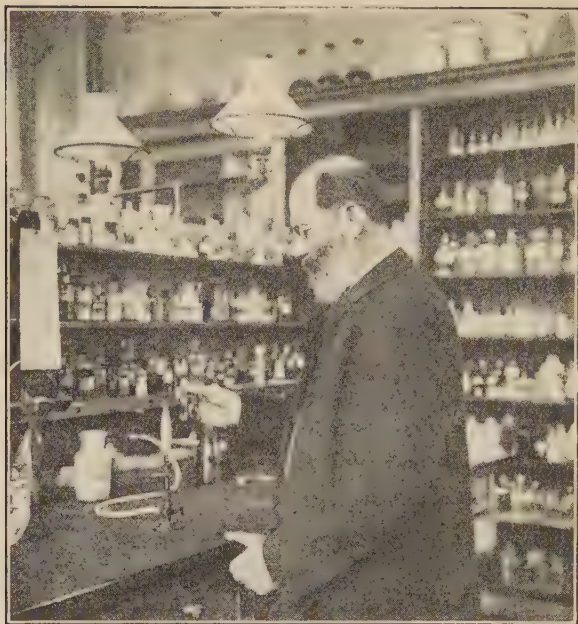


FIG. 13. A chemical laboratory in which the "scientific attitude" is illustrated.

achieved what it has because its chemists never attack a problem with a prejudice in favor of one view as opposed to another. "Give me the facts of the case," is their eternal cry. Professor James Harvey Robinson, the noted historian, has beautifully expressed this in the following sentence, which is worth reading more than once, and thinking over: "These men [the scientists, and therefore the chemists] crave a meticulous (painstaking) precision of observation, measurement, and statement quite alien to the other teachers of men. They exhibit an almost shocking insensibility to the cherished motives of belief. They do not ask whether what is sought is either right or wrong, beautiful or ugly, useful or futile, comforting or distressing. They only ask whether what is found is an instance of something really happening."

Robinson expresses in his sentence what has come to be known

as the *scientific attitude* or the scientific mood; and one of the most valuable reasons for studying chemistry, apart from making one familiar with some of the most glorious achievements of the race, is to discover just what this scientific mood or scientific attitude really is. For the citizen who takes his citizenship at all seriously, who is interested in the welfare of his country and the race, may well ask himself this question: If this scientific attitude — the critical attitude of examining without prejudice — has done so much in the field of chemistry (and of science), why should we not suppose that when applied to other fields of human endeavor, to our politics and to our social life, this scientific attitude would also prove of great value? So far, as your history and civics and English will teach you, the scientific attitude has been a stranger in fields other than chemistry, physics, biology, and the sciences based on these three.

We see, then, that an important reason for studying chemistry — the most important, many are convinced — is that, with the scientific attitude firmly fixed, it tends to make a better citizen of the individual. Further, we must remember that we live in an industrial era, every phase of which deals with a chemical problem; and to treat such problems intelligently is the duty of the citizen. And we cannot overlook that in the close connection which exists between chemistry and medicine, the factor of health and all that it involves in developing a sound race, is intimately bound up with a knowledge of chemistry.

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- E. E. Slosson, *Creative Chemistry* (The Century Co., 1919), Introduction and pp. 3-13.
C. H. Herty, etc., *The Future Independence and Progress of American Medicine in the Age of Chemistry* (Chemical Foundation Inc., 1923).
E. E. Slosson, "The Human Side of Chemistry," *Journal of Industrial and Engineering Chemistry*, vol. XIV (1922), p. 887.
J. H. Robinson, *The Mind in the Making* (Harper and Bros., 1923).

- J. H. Robinson, *Humanizing of Knowledge* (Doubleday Doran & Co., 1926).
Jesse E. Whitsit, "Science and Intellectual Integrity," *Bulletin of High Points*,
vol. VI, p. 12, and vol. VII, p. 3 (1925) (published by the Board of Educa-
tion, New York).
Benjamin Harrow, *Eminent Chemists of Our Time* (D. Van Nostrand Co., 1927).

QUESTIONS

1. Give several examples to show how chemistry affects your everyday life.
2. Name two substances which were formerly waste products and are now being utilized.
- 3. What are some of the characteristics of the ideal chemist?
4. How might the study of chemistry make a man a better citizen?

OPTIONAL QUESTIONS

5. Make a report on some recent outstanding achievement in chemistry. (See list of books under Supplementary Reading.)
- 6. Explain to what extent you use the scientific attitude in your debates.



FIG. 14. An alchemist in his laboratory.

CHAPTER II

HISTORICAL INTRODUCTION

7. The chemistry of the ancients. The science of chemistry is little more than a century old; but many substances classed as *chemical* have been known and manufactured during the prehistoric and ancient periods which ended with the civilizations of Egypt, Greece, and Rome. The ancients obtained salt by the evaporation of sea water with the help of the sun's rays; they found soda as a natural deposit in Egypt and used it as a cleaning agent and in the manufacture of glass; they discovered Fuller's earth, a white clay, which was also used as a cleansing agent before soap was manufactured; they were familiar with limestone and marble, which were employed as building materials, and which were also converted into lime by heat and used in the preparation of mortar; they knew that sand had to be mixed with soda to get glass; they prized the emerald, the topaz, and the diamond; they obtained wine and vinegar by fermenting the grape and other

fruits; they extracted turpentine and waxes from plants and oils and fats from the bodies of animals; they were familiar with gold, silver, and copper found in the native state; and they developed methods of extracting from their ores such metals as tin, iron, and lead.

8. The period of alchemy, extending from the early Christian era to about 1650 A.D., is noted as the attempt by forerunners of the chemists, the *alchemists*, to discover the "philosopher's stone," with the help of which all "base" metals, such as tin and lead, could be converted into gold, and the "elixir of life," with the use of which life could be prolonged indefinitely. Little real progress was made during this period, though many isolated chemical facts were accumulated. Towards the close of the sixteenth century many alchemists turned their attention to the preparation of various drugs and to their action on the human body. "Alchemy," writes Bacon, in his *Advancement of Learning*, "may be compared to the man who told his sons he had left them gold buried

somewhere in his vineyard; where they, by digging, found no gold, but by turning up the mould about the roots of the vines, procured a plentiful vintage. So the search and endeavors to make gold have brought many useful inventions and instructive experiments to light."



Keystone View Co.

FIG. 15. Lavoisier.

9. The modern period in chemistry was ushered in by *Lavoisier* (1743–94), who for the first time explained the true nature of burning. From then on chemistry made rapid strides; largely, as many believe, because of the introduction into

the science of quantitative methods of investigation. By "quantitative methods" we mean methods which make use of weighing and measuring. Without weighing and measuring, as experience has taught us, no branch of knowledge can become a science.

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- S. A. Arrhenius, *Chemistry in Modern Life* (D. Van Nostrand Co., 1925), pp. 1-16.
- M. M. Pattison Muir, *Heroes of Science, Chemists* (Society for Promoting Christian Knowledge, London, 1883), pp. 1-29.
- T. M. Lowry, *Historical Introduction to Chemistry* (The Macmillan Co., 1915), pp. 1-11.
- E. H. Smith, "Some Early Chemical Symbols," *Journal of Industrial and Engineering Chemistry*, vol. XVI (1924), p. 406.
- W. D. Richardson, "The Lost Arts of Chemistry," *Science*, April 7 (1911), p. 513.
- T. L. Davis, "The Emerald Table of Hermes Trismegistus," *Journal of Chemical Education*, August (1926), p. 863.
- William Foster, *The Romance of Chemistry* (The Century Co., 1927), pp. 3-17.
- E. R. Caley, "The Leyden Papyrus X," *Journal of Chemical Education*, vol. III (1926), p. 1149.

QUESTIONS

1. Name some of the substances found in nature by the Ancients.
2. Name some of the substances made by the Ancients.
3. What were the main objectives of many of the alchemists?

OPTIONAL QUESTIONS

4. Re-read paragraph 6 and then explain why chemists in one century have accomplished so much more than the alchemists in many centuries.
5. Prepare a report on the life and work of some one modern chemist.

CHAPTER III

THE EARTH. ELEMENTS

10. Earth, air, fire, and water. The ancients, headed by Aristotle, and the people of the Middle Ages, who followed Aristotle blindly, believed that out of earth, air, fire, and water everything in this world was formed. These four were considered

elementary bodies, in the sense that nothing simpler could be imagined; and elementary also in the sense that out of them complex bodies could be built. In our day these "elements" of the ancients, with the exception of fire, have all been shown to be susceptible of further analysis, of further simplification; which means that these elements are not elements at all in our sense of the word. In this chapter we shall devote ourselves to one of these elements, earth, and

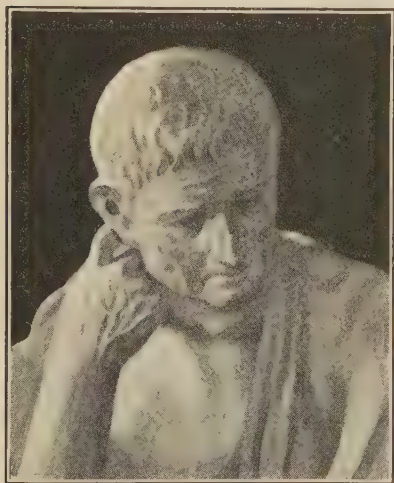


FIG. 16. Aristotle.

show how very far from a simple body this is; and in subsequent chapters we shall deal in an analogous manner with the other elements of the ancients.

11. The earth. It is believed that the earth was at one time part of the sun, as hot as is the sun; which means that at one time the earth was not a solid (as the portions of it which we can see

are to-day), nor a liquid, like the ocean, but a gas, like steam, or the invisible air which surrounds us. Something happened and a fragment from the sun detached itself. Then something else happened, and this detached fragment broke into two, a larger and a smaller portion. The larger portion, in the course of countless ages, became the earth and the smaller portion the moon.



FIG. 17. Crust of earth, showing strata of rock below the soil.

With this earth, which has been cooling ever since, and the interior of which we still believe to be very hot, go the oceans about us and the atmosphere above us; but as we shall discuss air and water in subsequent chapters, we shall at present confine ourselves to the solid portion of the earth — to the earth upon which we walk. And even here we know little beyond the surface of the earth. The deepest borings are mere scratches when compared to the depth of the earth. Our examination, then, confines itself to the crust of the earth. The earth is covered with soil, gravel or sand, the thickness of which may be a few inches or several hundred

feet. Sometimes rock is exposed to view ; but whether so or not, it is always found below the soil, gravel or sand, and extends, we suppose, many, many miles. The soil, the water, the mud underlying the ocean — which is formed when the large quantities of water on the surface of the earth fill depressions — and the rock constitute the crust of the earth. The soil supports plant life, and plant life, in turn, supports animal life, including man, either directly or indirectly. The rocks, in turn, supply us with our minerals: our coal, our limestone, iron, borax, rocksalt, copper, gold, silver, etc.

12. Elements. When the chemist comes to analyze the crust of the earth, he finds that it can be resolved into about 90 simple

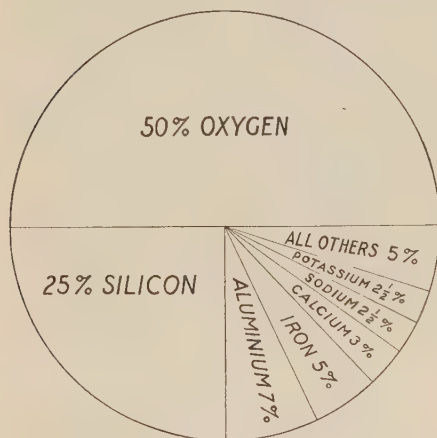


FIG. 18. Per cent of elements in the earth.

substances — substances which so far have defied further simplification ; and these he calls *elements*. For example, one of the substances found in the earth, sand, may be broken up into silicon and oxygen. Silicon and oxygen, however, have not been broken up into anything else. Therefore sand is not an element, whereas silicon and oxygen are elements. Out of these 90 elements, every-

thing that you can see — and this includes not only the earth, but the vegetable and animal kingdoms, including man — is built up. One of these elements, oxygen, constitutes about one-half of the solid material of the earth, one-fifth of the atmosphere, and about eight-ninths of the oceans. Next in abundance in the solid crust of the earth is the element silicon, which, combined with oxygen, forms sand. Aluminum, the third element, is the most abundant

of all metals, though it never is found in the free state. Then, in decreasing amounts follow such metals as iron, calcium, sodium, potassium, and magnesium — all elements, remember. Nitrogen, the predominant element in the atmosphere, and hydrogen, which forms one-ninth part by weight of water, compose but a very small portion of the solid earth. And in the earth are found, either combined or otherwise, such common elements as gold, silver, copper, lead, tin, zinc, and carbon.

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- H. A. Webb and J. J. Didcoct, *Early Steps in Science* (D. Appleton & Co., 1924), pp. 409-433.
W. H. Snyder, *Everyday Science* (Allyn and Bacon, 1919), pp. 247-276.
H. G. Wells, *Outline of History* (The Macmillan Co., 1920), chaps. i and ii.
F. W. Clarke, "The Data of Geochemistry" (Government Printing Office, Washington, 1920), pp. 1-41.

QUESTIONS

1. Why is earth no longer considered an element?
- 2. What are some of the more common elements of which the crust of the earth is made?
- 3. You will discover in a subsequent chapter that water can be decomposed into hydrogen and oxygen. Why do we no longer agree with the ancients that water is an element?

OPTIONAL QUESTION

- 4. Give a brief account of the possible origin of the earth upon which we live.

CHAPTER IV

THE AIR. OXYGEN AND NITROGEN

13. The ancients, we have already said, considered the air an element. As they were unable to resolve air into any simpler substances, they were justified in considering it an element. To-day we know better. The air consists of a number of elements.

14. Origin. We have already pointed out that in ages gone by the earth was probably part of the sun, and that something happened and part of the sun tore itself away. This part eventually, we believe, became what is now known as the earth. The earth was originally a very hot mass, probably so hot that everything was in a state of vapor, like steam. Since that time the earth has been cooling, and some of the vapor has changed to liquid, like the ocean, and some to solid, like the solid crust of the earth. But some of the contents of the earth have refused to become liquids or solids, and they have remained as gases encircling our globe. These gases we call the *air* or *atmosphere*.

15. The atmosphere. Most of the air is found near the surface of the earth. It becomes less and less dense the higher we ascend, as any airman or mountain-climber can tell you. How high up some of this air reaches we have no means of telling, but we do know that men in airplanes have reached a height of over five miles, and men in balloons a height of nearly nine miles. At such heights breathing becomes most difficult, suggesting at once that something in the air is necessary for life. This "something" is found plentifully on the surface, but in smaller and smaller amounts as we ascend higher and higher. This something we know now to be the gas *oxygen*, about which we shall have more to say presently.



FIG. 19. Air in motion—a cyclone.

16. Air has weight. The air cannot be seen, and yet we have many evidences of its presence. The breeze, the wind, the storm, the hurricane, tell you what it can do when set in motion. If you

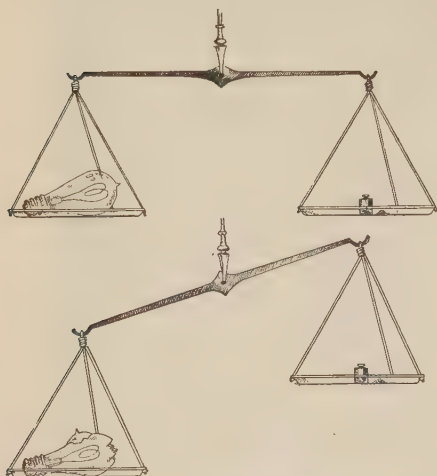


FIG. 20. Bulb before and after puncturing to illustrate that air has weight.

pump air into a tire of an automobile, you can show that it occupies space. If you weigh the tire before and after inflation, you will see that the tire has increased in weight after you have blown air into it; proving, thereby, that this air has weight. Or, if you wish, you can weigh an electric bulb before and after you have broken the tip. To what would the increase in weight be due, do you suppose?

17. What the air is made of. In 1789 Lavoisier, a famous French chemist, performed an experiment which threw much light on the composition of the air. He put some mercury into the retort, the neck of which was bent and connected with a bell-jar of air inverted over liquid (Fig. 21). He heated the retort and observed that the mercury turned into a red powder, and that the liquid in the bell jar rose a certain distance. Animals introduced into the air which remained in the bell jar died within a few moments "and lights were extinguished in it at once, as if they had been plunged into water." The red powder was next heated, and the

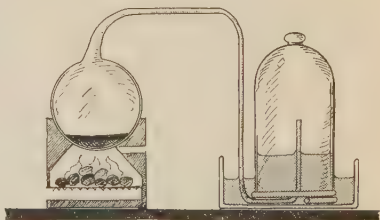


FIG. 21. Lavoisier's experiment to show the composition of air.

gas which was evolved was collected. A candle plunged into this gas "gave out a dazzling light; charcoal, instead of burning quietly as in ordinary air, burnt with a flame and a brightness of light which the eye could scarcely bear." And the gas "was much more fit" to support life than ordinary air.

How did Lavoisier explain his results? This is what he said: When I first heated my mercury in the retort, it combined with a part of the air (filling the apparatus) to form the red powder. The rise of the liquid in the bell-jar showed that some of the air had been removed. The air which remained in the bell-jar, and which could not support life, nor allow a candle to burn in it, Lavoisier called *azote*, from a Greek word meaning "without life." We use the word *nitrogen* where the French use *azote*. The red powder formed in the retort undoubtedly represented mercury in combination with a part of the air, for when this red powder was heated, mercury was left behind, and a gas was given off in which a candle burned very brightly. This gas Lavoisier called *oxygen*.



FIG. 22. No burning or life without air.

In this way air was shown to consist of at least two gases: nitrogen, which does not support life, and in which a candle does not burn; and oxygen, which does support life, and in which a candle burns very brightly. Since Lavoisier's day, and quite within our own times, argon, helium, and several other gases have been shown to be present in the air.

18. The amount of oxygen and nitrogen in the air. A very simple experiment will give you a rough idea of the amounts of oxygen and nitrogen in the air. If you first moisten a test tube, and then add some powdered iron so that the sides are covered, and if you next invert the test tube over a pan of water, you will notice that the water gradually rises in the tube until about one-

fifth of the length of the test tube is reached. The four-fifths that remain contain a gas in which a lighted candle is immediately extinguished.

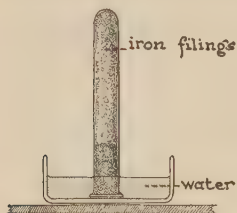


FIG. 23. Air contains about $\frac{1}{5}$ oxygen.

Obviously we are here dealing with the gas nitrogen, which is present in the air to the extent, roughly speaking, of eighty per cent. The one-fifth, or remaining twenty per cent, which was absorbed by the iron (to form iron rust), represents the oxygen. The iron combined with the oxygen of the air in the test tube just as

the mercury in Lavoisier's experiment combined with the oxygen of the air.

19. Carbon dioxide. You have perhaps performed the experiment of blowing into a test tube containing limewater and noticing how the latter turns milky. This milkiness is due to a gas, carbon dioxide, which we and all other animals exhale, and which finds its way into the air. The carbon dioxide which we find in the air also arises from the burning of coal and wood and the decay of vegetable and animal matter.

20. Carbon dioxide and oxygen cycle. Despite the huge quantities of carbon dioxide which find their way into the air, relatively small quantities are found there. This is due to the fact that plants take in carbon dioxide, which they need for building up plant foods and plant tissues. At the same time the plants give off oxygen, which very conveniently replenishes the oxygen consumed by us.

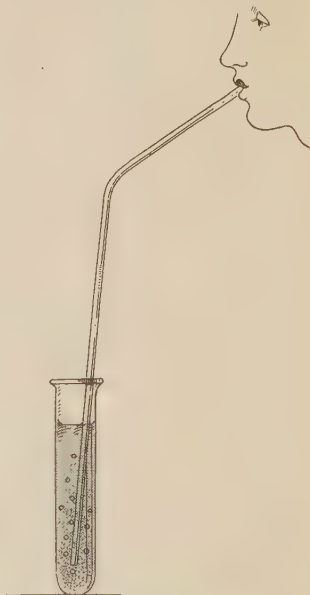


FIG. 24. Blowing into limewater. A test for carbon dioxide.

21. Dust. The air is not only filled with gases but with small solid particles to which the name of *dust* is given. Dust is really a mixture of many things — salts, limestone, clay, soot, products obtained from plants and animals, bacteria, etc. It becomes noticeable most readily when a beam of sunlight enters a darkened room; many minute particles are then seen in the track of the beam of light. But there are other evidences of its presence, such as the dust which accumulates on your clothes, and on the objects in the room. If you think this dust an outright nuisance, it might interest you to learn that without it no clouds or rain would be possible.

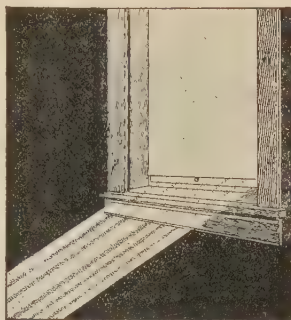


FIG. 25. Dust particles.

22. Moisture. Due to the constant evaporation of some of the waters about us, which is brought about by the heat from the sun, a certain amount of moisture is always found in the air. The actual amount of moisture which the air can hold at any one time depends upon the temperature of the air. After the air becomes saturated with moisture some of it may deposit in the form of rain or snow.

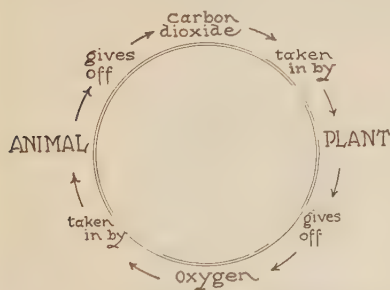


FIG. 26. The carbon dioxide and oxygen cycle.

The amount of water vapor in the air has a profound effect upon our comfort. If there is too much moisture in the air, evaporation from the body is retarded, giving rise to the sticky feeling so familiar during summer months; on the other hand, if there is too little moisture, we feel dry and uncomfortable. This is well illustrated in rooms heated by hot-air furnaces (Fig. 27). Here the discomfort may be alleviated by allowing the air to pass over a pan of water before it enters the room.

23. Air a mixture. The principal components of air, we have seen, are oxygen and nitrogen. These are not chemically com-

bined but exist quite free and independent of one another.

You will see when we come to study the properties of water that that substance consists of oxygen and hydrogen; but in water the oxygen and hydrogen are so united that you can no longer recognize either the one or the other. The two elements have combined, we say, to form

a *compound*. But the oxygen and nitrogen in the air have not combined; we can still recognize the oxygen and the nitrogen; just as, for example, it might be possible to recognize iron and lead powders when these two are mixed together. That is why we call air a *mixture*. There are a number of other very good reasons for calling air a mixture. One of the best of these is that the composition of air varies with time and place, whereas the composition of a compound is always the same.

24. Air can be liquefied. Theoretically all matter can exist in one of three

conditions, or, as we call them, states — solid, liquid, or gas. The commonest illustration of this is ice, water, and steam, all three

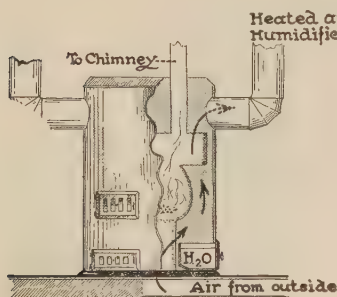


FIG. 27. Hot-air furnace.



FIG. 28. Air is a mixture.

of which, in reality, represent one substance in different states. We suppose that the essential difference between water and ice is that in the former the small particles composing it — the *molecules* — are further removed from one another than in the latter; and again, the essential difference between steam and water is that in the former the molecules are still further removed from one another; so far removed, in fact, that they tend to fly in all directions; and that is why steam, or any other gas, will fill a room so quickly. Some substances, like water, are liquid at the temperature common to the earth's atmosphere; other substances, like air, are gases at that temperature. Now it has been possible to cool air to such a point that it turns liquid; and this liquid air is largely used in the industries as a source of pure oxygen or nitrogen.



FIG. 29. A hot, sultry day.

25. Ventilation. A temperature of about 68 degrees Fahrenheit, with a fair circulation of air, but not one which produces a chilling draft, and a fair content of moisture, distinguishes "good" air from "bad" air. It was supposed at one time that bad air was bad because a lot of carbon dioxide accumulated in a room, whereas very little oxygen remained. This is no longer believed. Experiments on medical students have shown that you can actually increase the carbon dioxide content of a room to over 100 times the normal without producing any ill effects; provided, however, there is a good circulation. But, of course, this does not mean

that stale air is to be preferred to fresh air; it merely shows the comparatively harmless effects of inhaling large quantities of carbon dioxide, provided enough oxygen is present.

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QUESTIONS

- 1. What is the probable origin of the air which surrounds us?
- 2. How could you prove that the air has weight?
- 3. Describe Lavoisier's experiment to prove that air is a mixture of at least two gases.
- 4. Describe an experiment which will indicate the approximate amounts of oxygen and nitrogen in the air.
- 5. What are some of the other constituents in the air besides oxygen and nitrogen? What effects have they on our well-being?
- 6. Draw a simple diagram to show the interrelation of plants and animals.
- 7. Give at least two reasons for the statement that air is a mixture rather than an element.
- 8. What would you say are the requirements of a well-ventilated room?
- 9. Mention some good and some bad effects of dust in the atmosphere.

OPTIONAL QUESTION

10. Suggest a method for obtaining carbon dioxide in the three states.

CHAPTER V

FIRE. OXYGEN

26. Throughout the ages nothing, perhaps, has inspired greater wonder mixed with fear than fire. The ancients worshiped fire as a god when they failed to understand what it was; and as a matter of fact the idea underlying fire, burning, combustion — words closely related — was not understood until oxygen was discovered. From 1650 until well into the eighteenth century there was a popular theory that a substance burned because it contained a mysterious something called *phlogiston*. Substances which burned easily, like paper, for example, contained much phlogiston; substances like iron, which burned with difficulty, contained little phlogiston. Whenever something burned it lost its phlogiston. This theory received its death-blow from the French chemist, Lavoisier, who showed that substances after burning weighed more than before burning; whereas according to the phlogiston theory, they should, if anything, weigh less after burning. The overthrow of the phlogiston theory was also brought about by Priestley's discovery of oxygen and by Lavoisier's extensive experiments on the nature of combustion, in which he showed that what we call *burning* is a chemical reaction in which a substance combines with oxygen in such a manner as to give off noticeable amounts of light and heat. In other words, the common forms of burning or combustion take place only because oxygen is present. In every case of burning the substance burned undergoes a complete change — a change so complete that it can no longer be recognized. This is what we mean by a *chemical change*. On the other hand, when we stretch a rubber band or break a piece of glass, in spite of these

changes, we can still recognize both rubber and glass. This is what we mean by a *physical change*.

27. Oxygen. We have already seen in the previous chapter that one of the gases in the air causes a candle to give out a dazzling light; it strongly *supports combustion*, as we would say to-day. The actual discovery of oxygen is attributed to Joseph Priestley, an Englishman who spent his last few years in America. He was

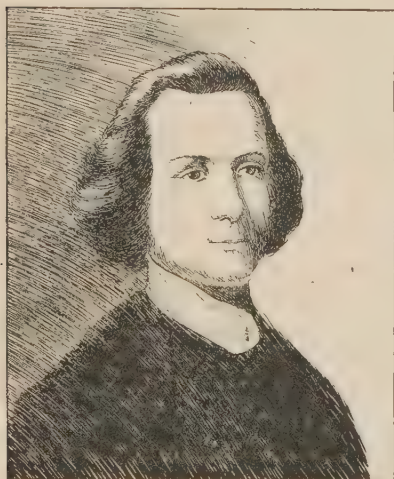


FIG. 30. Priestley.

interested in discovering what kinds of gases were formed when various substances were heated. One of these substances was mercuric oxide, a red powder, which, as we have already seen in the last chapter, Lavoisier used to prove some facts regarding the composition of the air. In August, 1774, some three years before Lavoisier's experiments, Priestley heated his mercuric oxide and obtained a gas in which "a candle burned with a remarkably vigorous flame . . . and a

piece of red-hot wood sparkled in it . . . and it consumed very fast." He placed a mouse in the gas, and the mouse lived on and on. He breathed it himself and his "breast felt peculiarly light and easy for some time afterwards."

28. Oxygen occurs in great abundance. We have already seen that about one-fifth of the air we breathe is oxygen, oxygen in the free state. We shall see in the next chapter that in the combined state — in a state in which it cannot be recognized — oxygen makes up 89 per cent of water. In a combined state it is also present in large quantities in our bodies and in the earth's crust.

29. Methods of obtaining oxygen. One way has already been described; it is the method used by Priestley and by Lavoisier: heating mercuric oxide. The red powder in the test tube (Fig. 32) is heated. A glowing splinter inserted into the tube bursts into

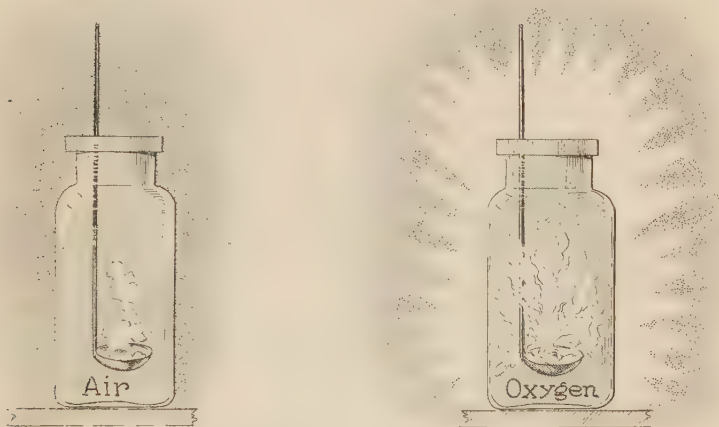
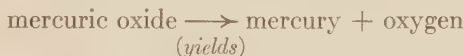


FIG. 31. Phosphorus burning in air as compared to phosphorus burning in oxygen.

flame, indicating the production of oxygen. An inspection of the sides of the tube will reveal small globules of mercury. The reaction can be summarized by stating that



Mercuric oxide consists of the elements mercury and oxygen in chemical combination. This method is too expensive for common use.

Another way, and a more common one, is to heat potassium chlorate, a white, crystalline substance, and collect the gas over water. By the addition of manganese dioxide to the potassium chlorate, the oxygen is given off more rapidly and at a lower temperature. Careful examination will show that though the potassium chlorate has changed as a result of the heating, the

manganese dioxide has not. Not only has the manganese dioxide not changed, but the same amount is found at the end of the reaction as in the beginning. In subsequent chapters we shall repeatedly come across substances which, like manganese dioxide, accelerate chemical reactions without themselves undergoing any

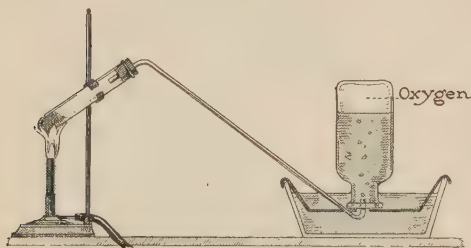


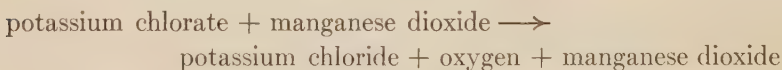
FIG. 32. The preparation of oxygen by the decomposition of mercuric oxide or of potassium chlorate.

permanent change; such substances are known as *catalytic agents*.

At this point one or two analogies may be of help in giving you an insight into a catalytic agent, but remember that the analogy, at best, is but a crude

representation. A catalytic agent has been compared to the oil used in a piece of machinery in order to make the parts move more smoothly. A cruder, though perhaps more effective, analogy is to suppose the teacher to be a catalytic agent. When he is out of the room, there is usually much commotion. Things quiet down immediately upon his return. The change has been brought about by his mere presence, and he himself has undergone no change in the process.

The reactions that take place with potassium chlorate and manganese dioxide may be summarized thus:



Or, since the manganese dioxide does not change in any way, it is customary to omit this substance from the statement and merely write



This is a convenient method of summarizing a chemical reaction and is called a *word equation*. Potassium chlorate consists of the

elements potassium, chlorine, and oxygen, all three chemically combined. When the substance is heated, it loses its oxygen and only the potassium and chlorine remain in chemical combination. This compound of potassium and chlorine is called potassium chloride.

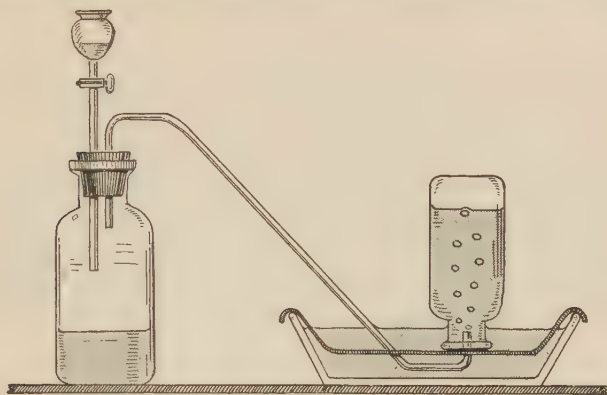


FIG. 33. The action of sodium peroxide on water.

One of the most convenient laboratory methods of preparing oxygen is to allow water to react with sodium peroxide. By merely bringing the water and the sodium peroxide in contact with one another, oxygen is liberated.

Commercially, oxygen is obtained either by the decomposition of water with the help of an electric current or from liquid air. If liquid air is exposed, the nitrogen, which evaporates more readily than the oxygen, escapes first, leaving behind almost pure oxygen.

30. Oxidation. We have seen in a previous chapter that when iron rusts it combines with oxygen, among other things. The combination of a substance with oxygen is called *oxidation*. In our bodies the carbon and the hydrogen in the foods we eat combine with oxygen. We say that the carbon and the hydrogen are oxidized and the process is called oxidation.



FIG. 34. A lighted match.
Rapid oxidation.

31. Oxides. The combination of an element with oxygen gives rise to an *oxide*; thus carbon, hydrogen, iron, magnesium, phosphorus, and many other elements combine with oxygen, in each case forming an oxide. Oxides are quite common in nature. We find such oxides as those of copper, iron, lead, aluminum, zinc. Can you give a reason for this?

32. Properties of oxygen. The mere fact that oxygen is a constituent of the air, and that the air is invisible, means that oxygen is a colorless gas. Neither has it any taste or odor. Since the air covers our oceans and rivers, and the amount of oxygen in the air does not change, this would also indicate that the gas is not very soluble in water. Oxygen is but slightly soluble in water; and that is why we can collect the gas over water. Chemically speaking, oxygen is considered an *active element*, because it combines so readily with other substances. Recall the fact that oxygen combines readily with many elements to form oxides. It unites with hydrogen to form hydrogen oxide, which is commonly called water. Compared with oxygen, nitrogen, the other constituent in the atmosphere, is an inactive element, since it is quite a difficult task to get it to combine with other substances.

What are called the *physical properties* of gases are, among others, their color, taste, odor, solubility, and specific gravity.

33. Combustion and burning. Strictly speaking, any chemical reaction in which heat and light are evolved is one of combustion or burning. As a rule, however, combustion refers to the union of some substance with oxygen, with the evolution of heat and light.

34. Spontaneous combustion. Different substances will begin to burn at different temperatures. Paper burns at a lower temperature than coal. But a thin strip of wood, as you well know, will burn much more readily than a thick log of wood; and if the log of wood be cut into thin strips, the strips will burn more readily than the original log. The actual amount of material exposed to

the air, or the actual amount of material that comes into intimate contact with the oxygen in the air, is all important. Obviously, more of the material in the thin strips of wood comes in contact with the oxygen than in the original thick log. The lowest temperature at which a substance will burn is spoken of as the *kindling temperature*.

A substance like iron will rust when exposed to the air. The iron, among other things, combines with the oxygen of the air.

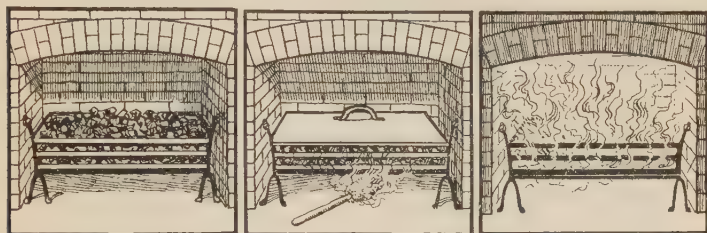


FIG. 35. Three conditions are necessary for combustion. In the first drawing the kindling temperature has not been reached; in the second oxygen is excluded; in the third all three conditions are satisfied.

Oxidation occurs and heat is generated. But as such an oxidation is a slow process, we call it *slow oxidation*; and of course the amount of heat given off in a short time is quite small. Another example of slow oxidation is the drying of linseed oil upon exposure to the air. This linseed oil finds extensive use in paints; and sometimes oily rags, full of linseed oil, are carelessly thrown into a cupboard, where the ventilation is poor, and where, in consequence, the heat produced in the course of oxidation does not escape. The heat accumulates and accumulates until the kindling temperature of the material is reached, and then there is a sudden burst of flame. The generation of fire under such conditions is known as *spontaneous combustion*.

The three conditions necessary for combustion are, therefore, *combustible material* (fuel), *a supporter of combustion* (oxygen), and *kindling temperature*. The absence of any one makes burning impossible.

35. Uses of oxygen. Without oxygen life would be impossible. Just as a piece of coal when burning combines with the oxygen, so

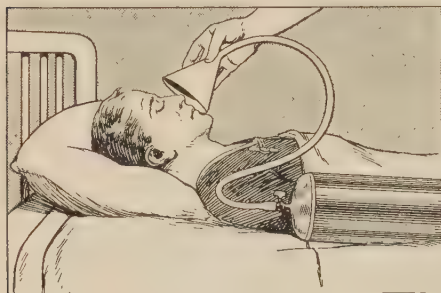
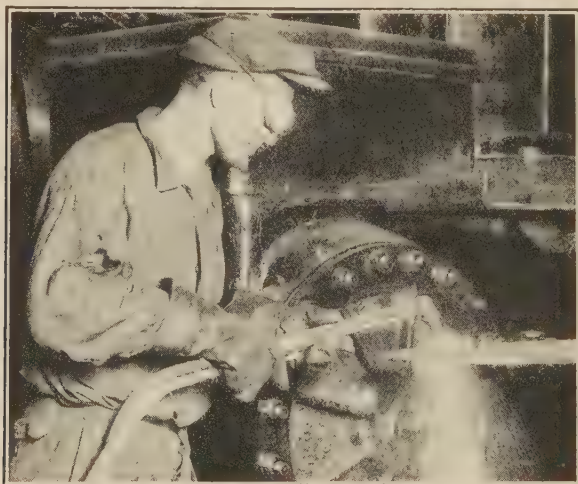


FIG. 36. Oxygen administered to a pneumonia patient.

our foods when oxidized in the body combine with oxygen. This process of oxidation in the body is essential to life. That is why whenever oxidation, or respiration, becomes impeded (as in such a disease as pneumonia) oxygen is often administered.

Oxygen also finds uses in the industries. The oxy-acetylene blowpipe is a good example. Here oxygen and acetylene are burned together, yielding an ex-



Ewing Galloway

FIG. 37. The oxy-acetylene blowpipe.

tremely high temperature — so high, in fact, that one can cut manholes in steel plates. This has largely displaced the oxy-

hydrogen blowpipe, which does not develop so high a temperature. Often the carbon which accumulates in automobile cylinders is removed by burning it with oxygen. Airplanes flying



Keystone View Company

FIG. 38. Working in mines by aid of tanks of oxygen.

at a high altitude and submarines are supplied with tanks of oxygen. Rescue parties in coal mines are equipped with such tanks.

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QUESTIONS

1. What is the difference in composition between a chlorate and chloride? Judging by analogy, what is the difference in composition between a sulphate and a sulphide, nitrate and nitride?
2. One jar contains oxygen and another air; identify each of them.
3. Why does a fire burn more vigorously when air is blown through it? What could be substituted for air to make the fire burn even more brightly?
4. If blowing air through a fire increases burning, why does blowing at a lighted match put it out? (Recall the three conditions for combustion.)

5. Define combustion, slow oxidation, kindling point, spontaneous combustion, and give an example in each instance.
6. Of the various methods suggested for the preparation of oxygen, which would you select for preparing a small quantity in the laboratory and a large quantity for commercial use? Give reasons.
7. Name the products formed when the following elements are burned in air: carbon, hydrogen, sulphur, phosphorus, zinc, iron, aluminum, and copper. Write the word equations.
8. Suppose phosphorus were heated in a closed vessel from which all the air had been exhausted, what would happen? Explain.
9. A forest fire is extinguished by back-firing. The fire department sometimes dynamites buildings. Explain a similarity between the two methods used.
10. Why is it necessary to aerate the water of an aquarium?
11. Foods contain the elements carbon and hydrogen. What do you suppose happens to these elements in the body? Coal contains carbon and hydrogen. What products are formed when coal burns?
12. Describe an experiment which will indicate the percentage of oxygen in the air.
13. Explain why wood cannot be set on fire while wet.
14. Paints dry rapidly when metallic oxides are added. Explain.
15. Explain the high temperature obtained by the use of the oxy-acetylene blowpipe.
16. Explain how water extinguishes a fire.
17. What is meant by "smothering a fire"?

OPTIONAL QUESTIONS

18. Outline an experiment to prove that the manganese dioxide does not change in weight when used in the preparation of oxygen.
19. Report on the work of either Priestley or Lavoisier in connection with the discovery of oxygen and in connection with our ideas of oxidation.
20. We breathe air which contains nitrogen. What happens to this gas in the body?
21. Explain how a fire may arise in a coal mine, the air of which is full of dust.
22. Name some ways by which rusting may be prevented.
23. When a furnace fire is very low, opening the damper wide may put out the fire. Explain.

CHAPTER VI

WATER. HYDROGEN

36. One theory, we have seen, has it that the earth upon which we live was at one time part of our sun, perhaps, and that in some way which we do not understand, tore itself away into space, becoming cooler and cooler as the years went by. Some of the elements combined and formed the solid crust of the earth; others did not combine but remained as gases; still others combined to form water, which surrounds us in the form of rivers and oceans, and which makes up a very large percentage of everything that has life, from the simplest plant to man himself.

37. **The composition of water.** In the laboratory you will be shown an apparatus somewhat like the one in the drawing, which is used to determine just what water is made up of. When we pass an electric current through the water (provided a little sulphuric acid is first added to it),

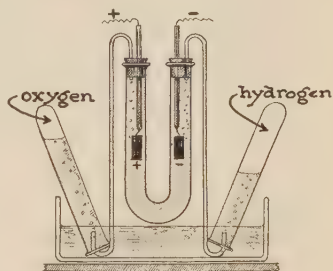


FIG. 39. Electrolysis of water.

gas collects in the two tubes, and twice as rapidly in the one as in the other. We can collect these gases, and we find that if we apply a glowing splinter to a test-tube full of one of them, the splinter will burst into flame, suggesting at once that this gas is none other than *oxygen*; on the other hand, if we apply a lighted match to the other tube, the gas in the tube catches fire, burns with a pale blue flame, and if air is mixed with the gas, a slight “pop” or explosion will take place. All this, as you will presently discover, is characteristic of a gas which we call *hydrogen*.

Now what does this experiment on the electrolysis of water show you? Obviously, that water can be resolved into two gases, into two elements, oxygen and hydrogen; and, furthermore, at any one time, twice as much hydrogen is formed as is oxygen. For you will discover that it is the gas which burns with a pale blue flame which at any one time occupies twice as much space as the gas which causes a glowing splinter to burst into flame. You learn, then, that in terms of the amount of space occupied, or, as we like to put it, in terms of volume, the hydrogen is to oxygen (in water) as 2 is to 1. Now the very remarkable thing is that you can take water from any source you like, whether from a lake, or from rain, or from a river, put it into our apparatus for electrolysis, pass a current through the water, and you will always obtain hydrogen and oxygen in the proportion by volume of 2 to 1.

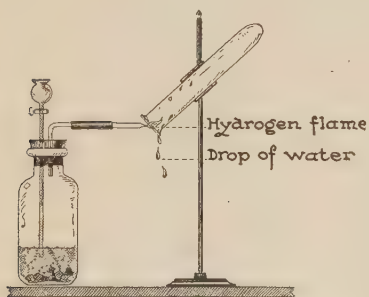


FIG. 40. Burning hydrogen to form water.

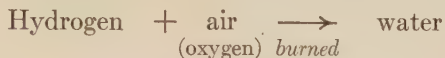
Water, we see then, consists of at least two elements, oxygen and hydrogen; but apparently so combined that in the liquid you cannot recognize either the one element or the other — not until you have broken the union asunder, by the electric current, for example, as we have done. Water, you see, is a compound. But,

you may say, how do we know that water consists of nothing else? We can prove this in a very simple way. We will show you, later on, that not only can we decompose water:



but we can cause oxygen and hydrogen to reunite to form water again. For example, you will learn that when hydrogen is burned, water is actually formed. Now what do we mean by “burning”? We have already learned in the last chapter that it is essentially the combination of a substance with oxygen. In other

words, then, when hydrogen burns, it combines with the oxygen of the air to form water. This, incidentally, also proves that water really consists of nothing else but the two elements oxygen and hydrogen; and the burning of hydrogen may be represented by the equation



Notice that the union of oxygen and hydrogen to form water is a case of *synthesis*, or the building up of a compound from its elements. Notice, further, that the electrolysis of water to form oxygen and hydrogen is a case of *analysis*; or the decomposition of a compound into simpler units.

38. Occurrence and properties of water. Water is everywhere, not only in the form of rivers and oceans and rain, where it can be

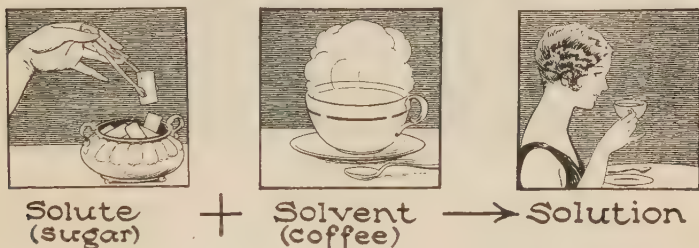


FIG. 41.

recognized readily enough, but as part of our bodies, as part of our foods, as part of the very rocks, even, of the earth's crust. It is the best *solvent* we have, meaning by solvent a "dissolver of things"; for nothing, not even the most resistant rocks, is able to withstand its solvent action. When this water dissolves a substance, such as salt or sugar, say, the salt or sugar is then spoken of as the *solute*, or the substance which is dissolved, and the solvent and the solute together are spoken of as a *solution*. If there is very little of the solute, then we have a *dilute solution*. If we have much of the solute, then we have a *concentrated solution*.

If we take a half tumbler of water and add sugar (a small amount at a time, stirring constantly) until no more of it dissolves, then we get a *saturated solution*. But, this saturated solution holds good for just one particular temperature; for you know from your experience that more sugar will dissolve in hot than in cold water; and, as a rule — for there are exceptions — solubility varies directly with the temperature; which is a compact way of saying that the hotter the water the more readily will a substance dissolve in it.

Under ordinary conditions water is an excellent example of what we call a *liquid*. But in the winter water may freeze and we get ice, which is a *solid* form of water, and we can heat water until it boils, and we get steam, which is water in the *gas* form.

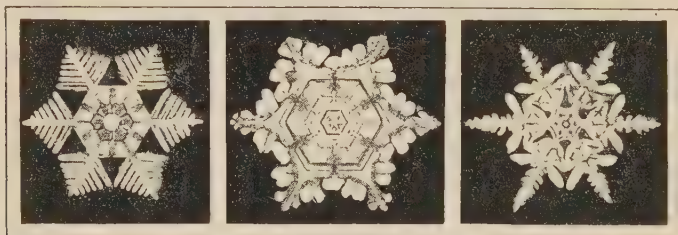


FIG. 42. Snow crystals.

39. Crystals. A substance dissolved in water will often *crystallize* from solution when the excess water has been evaporated. The substance often assumes a very definite geometrical shape — cubic, hexagonal, etc. The crystals so formed are due to the combination of the substance with a definite quantity of the water, this water being commonly referred to as *water of crystallization* or hydration. These combinations of substances with water are often referred to as *hydrates*.

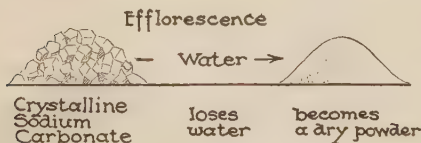


FIG. 43. Efflorescence.

40. Efflorescence and deliquescence. Crystals of washing soda become powdery and opaque on standing. This is due to the fact that

they lose their water of hydration. Substances, such as washing soda, which lose their water of crystallization on standing are spoken of as *efflorescent*. Other substances, like calcium chloride, become moist on exposure; these take in water, and are known as *deliquescent*.

It is commonly considered that a deliquescent substance is one which actually gets wet and melts, and that when a substance absorbs moisture without apparently getting wet, such as wool, we call it *hygroscopic*.

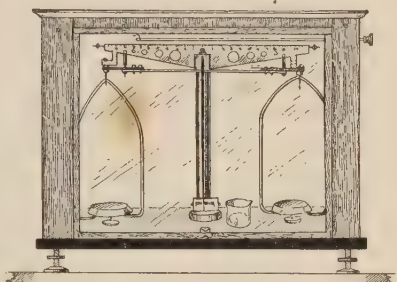


FIG. 44. Calcium chloride used in balance case to keep the air dry.

41. Impurities in water. We have pointed out that water is a great solvent. It contains all kinds of substances—mineral matter, which has no life, and organic matter, some of which may have life. The mineral matter, as a rule, is not harmful if the water is to be used for drinking purposes, but may be quite harmful if the water is to be used for washing and industrial purposes.

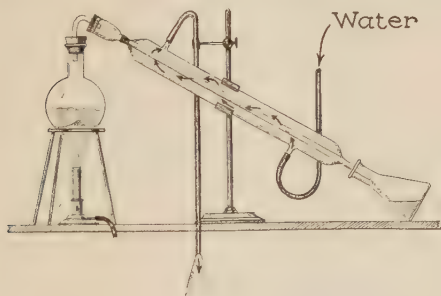


FIG. 45. Purifying water by distillation.

The organic matter, on the other hand, which very often includes germs of all sorts, may prove quite dangerous to health and not harmful in the industries. The simplest method of purifying a water is to distil it, which is simply a method by which the water is heated

until it boils, and the vapor which is given off is cooled and collected. The heat kills any harmful germs, and the various substances dissolved in the water remain in the flask. When

water is needed in very large quantities, or when it has to be supplied to the people in a town, purifying water by distillation is out of the question; the cost for the necessary apparatus and the fuel required to heat the water would be enormous. In such a case, the water is purified by a double process of filtration and chemical treatment. The filtration usually consists in first forming a fine suspension in the water (a "jelly-like precipitate" is what it is usually called) by adding alum or alum and lime, allowing this suspension to settle, which by so doing drags down the suspended impurities in the water and most of the germs, and finally filtering the water through a bed of sand and gravel. To make certain that all harmful germs are destroyed, the water is treated with small quantities of chlorine before being sent through pipes into the houses.

42. Hydrogen. In this chapter we learned that water is composed of oxygen and hydrogen chemically combined. We have already (par. 32) become familiar with the properties of oxygen; now we must study some of the properties which characterize hydrogen.

43. Preparation of hydrogen. We have seen how hydrogen (and oxygen) can be prepared by the electrolysis of water.

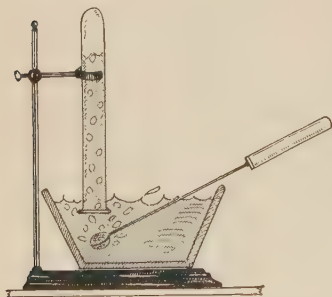
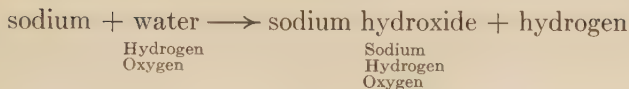


FIG. 46. The action of sodium on water.

This experiment shows, among other things, that hydrogen is a constituent of water. Now certain metals, such as sodium or potassium, will react violently with water, giving off a gas which burns with a pale blue flame, and which is really hydrogen. Since neither the potassium nor the sodium contains any hydrogen, we must assume that the hydrogen comes from the water; and here we

have additional evidence that one of the components in water is hydrogen. The reaction can be expressed by the equation



Sodium and potassium react with water just as soon as they come in contact with it. Other metals do not react so readily with water. With mag-

nesium and aluminum, for example, the hydrogen will be given off at a very slow rate indeed unless the water is first heated. With iron even more strenuous measures have to be taken: the

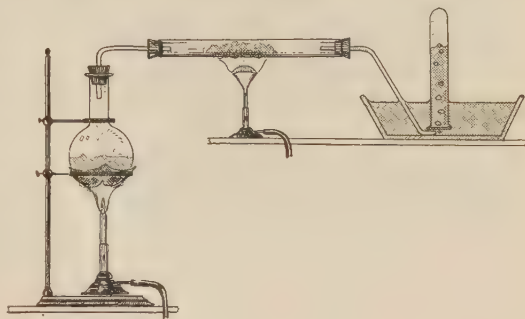
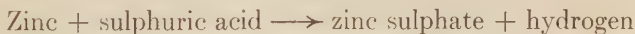


FIG. 47. Hydrogen made by passing steam over red hot iron.

metal is heated very strongly in a tube and steam passed through the tube.

A common method of preparing hydrogen, and one which you will use in the laboratory, is to bring a metal, such as zinc, and an acid, such as sulphuric acid, together:



Hydrogen
Sulphur
Oxygen

Zinc
Sulphur
Oxygen

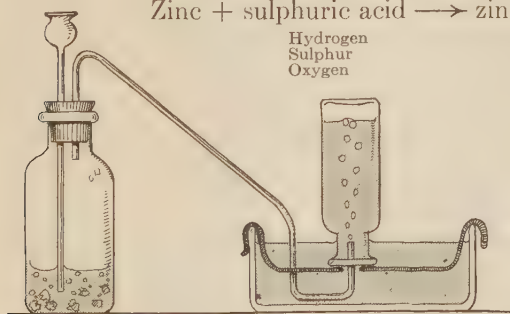


FIG. 48. The action of zinc on sulphuric acid.

other acids, like hydrochloric acid or acetic acid, may be substituted for sulphuric acid.

44. Metals and their activity. The accompanying table of elements, in the order given, is worth becoming familiar with.

Potassium
Sodium
Calcium
Magnesium
Aluminum
Zinc
Iron
Tin
Lead
Hydrogen
Copper
Mercury
Silver
Platinum
Gold

You will notice that metals like potassium and sodium, which react so readily with water, are near the top; that metals like magnesium and aluminum, which do not act quite so readily, are a little further down; and that a metal like iron, which has to be heated pretty strongly before it will combine with water, is still further down the list. What would you say this grouping of metals signifies? That those near the top are more active than those further down in the list. In fact, the order of these elements is in the order of their decreasing chemical activity, potassium being more active than sodium, sodium more active than calcium, and so on. But this reactivity is not merely confined to the action of a metal on water; it holds equally well when we consider the action of a metal with acids. If zinc reacts readily with sulphuric acid, potassium or sodium reacts even more readily with this acid.

Now why do you suppose that hydrogen is found in this list — hydrogen which is not a metal at all? Merely to inform you at this point that only those metals above hydrogen react at all with water or acid, whereas those below hydrogen do not react with water or acid — at least, not appreciably. Notice, for example, gold, the last on the list. You have no hesitation in washing your gold ring with soap and water; nor has the jeweler any hesitation in testing his gold with nitric acid.

45. Properties of hydrogen. You will have discovered from experiments in the laboratory that, like oxygen and air, hydrogen is a colorless gas without any odor. Since it is collected over water, it, like oxygen, is at best but slightly soluble in water. It is the lightest among the gases, and hence was used extensively for filling balloons and airships. The objection to the use of hydrogen

in airships is that it is so very readily inflammable, for you will remember that hydrogen burns easily, forming water; hence the substitution, since the war, of helium in the place of hydrogen; for helium, though somewhat heavier than hydrogen, is still much lighter than air and, above all, does not burn. You will learn that helium is one of the "inert" gases; it does not combine with any element; which, of course, means that it does not combine with oxygen, among other things.



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FIG. 49. Airship.

Hydrogen very readily abstracts oxygen from other substances. For example, if we pass hydrogen into a tube containing copper oxide — which is composed of copper and oxygen — and we heat it, the black copper oxide is changed into a reddish deposit, which we can prove to be copper, and moisture is formed in the cooler portion of the tube. The reaction can be represented as

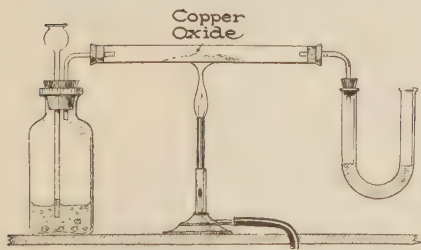


FIG. 50. The reaction between copper oxide and hydrogen.

Now water, as you of course know, is composed of oxygen and hydrogen. So that what has really happened in this reaction is that hydrogen has dragged away the oxygen from the copper oxide to form water. This property that hydrogen possesses of abstracting oxygen from other substances makes it a *reducing agent*. At this stage let us emphasize one or two points. The copper oxide in this reaction parts with its oxygen; hence the

copper oxide is an *oxidizing agent*. The hydrogen takes the oxygen; the hydrogen is therefore a *reducing agent*. The copper

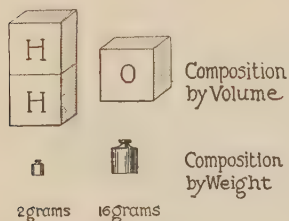


FIG. 51. The volume and weight relationships of water.

oxide is *reduced* to copper; the hydrogen is *oxidized* to water. You will find that wherever there is oxidation there is also a corresponding reduction; and vice versa.

Whereas the volume relationship of hydrogen and oxygen in water is as 2 is to 1, the weight relationship is quite different. To obtain this weight relationship, we pass hydrogen through heated copper oxide, as shown in the figure. The reaction may be represented:



You will notice that the hydrogen takes the oxygen from the copper oxide to form water. If, now, we weigh the copper oxide before and after the experiment, the difference in the two weights will give us the weight of oxygen used to form water. If, again, we collect the water and weigh it, the difference between it and the weight of the oxygen used to form water will give us the weight of the hydrogen. An actual experiment will show that the oxygen weighs eight times as much as the hydrogen.

To summarize these relationships, though there is twice as much hydrogen as oxygen by volume, owing to the fact that the oxygen is very much denser, by weight there is 8 times as much oxygen as hydrogen.

46. Uses of hydrogen. We have mentioned under oxygen (par. 35) the use of the oxy-hydrogen blowpipe for welding and cutting metals. We may also say that much hydrogen is now used

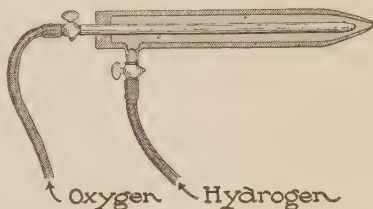


FIG. 52. The oxy-hydrogen blowpipe.

in the preparation of lard substitutes. The process is known as *hydrogenation* and consists in heating the oil and hydrogen in the presence of a catalytic agent like nickel. Another very important use for hydrogen, developed since the war, is in the manufacture of ammonia by the direct combination of hydrogen and nitrogen in the presence of a suitable catalyst. An even more recent development is the preparation of wood alcohol by the combination, under suitable conditions, of hydrogen and carbon monoxide. Lately a manufacturing process for making hydrochloric acid has been developed which involves the direct combination of hydrogen and chlorine. We have already referred to the use of hydrogen in balloons.



FIG. 53. A product of the hydrogenation of oils.

SUPPLEMENTARY READING

- William Foster, *The Romance of Chemistry* (D. Van Nostrand Co., 1927), chap. vii (Hydrogen and Water).
- H. L. Barnitz, "Hydrogenation of Oils," *Chemical and Metallurgical Engineering*, vol. XXII (1920), p. 745.
- D. H. Killefer, "Purifying Industrial Water," *Industrial and Engineering Chemistry*, vol. XV (1923), p. 915.

QUESTIONS

1. If a mixture of ten parts by volume of hydrogen and one part by volume of oxygen is ignited, name the product formed. What would be left over, and how much?
2. If a mixture of ten parts by weight of oxygen and one part by weight of hydrogen are ignited, what would be left over, and how much?
3. If a piece of iron pipe is left in a damp cellar, what compound would be formed?
4. What changes will take place when a hot saturated solution of sodium nitrate cools?
5. Give an example taken from your daily experience of a solution, a solute, a solvent, a crystal, an efflorescent substance, and a deliquescent substance.

6. How would you purify a city's water supply containing finely divided solid matter and bacteria?

7. When zinc, silver, sodium, and aluminum come in contact with water, what products, if any, are formed?

8. Write a word equation to show what happens when hydrogen is passed over heated iron oxide.

9. Write the word equations for the following: calcium plus water; iron plus sulphuric acid; zinc plus hydrochloric acid.

10. Complete the following sentences: vinegar is a — solution of acetic acid; syrup is a — solution of sugar.

11. Name a substance used in the home which contains a large percentage of hydrogen.

12. Write word equations illustrating (a) synthesis, (b) analysis.

13. "Water, water everywhere." Name several substances which contain an appreciable amount of water whose presence is not readily apparent.

14. Given the metals silver, copper, and aluminum, predict which one would react most readily with acetic acid.

15. Why is hydrogen used in airships? What are the objections to its use? Why is helium sometimes used in place of hydrogen?

16. What compound would you expect to be formed when hydrogen reduces lead oxide?

17. Paper, wood, and coal all contain hydrogen. Name one product which is formed when these substances are burned.

OPTIONAL QUESTIONS

18. If a mixture of ten parts by volume of hydrogen and three parts by volume of oxygen is ignited, name the product formed. What would be left over, and how much?

19. Referring to paragraph 45, calculate the percentage composition of water by weight from the following data:

Weight of calcium chloride and tube after experiment = 107.8 g.

Weight of calcium chloride and tube before experiment = 106 g.

Weight of copper oxide and tube before heating = 151.6 g.

Weight of copper oxide and tube after heating = 150 g.

CHAPTER VII

WHAT IS MATTER? THE ELECTRON THEORY

47. Many years ago Sir Isaac Newton, one of the greatest of all scientists, had written : “ It seems probable to me that God, in the beginning, formed matter in solid, massy, hard, impenetrable, movable particles ; of such sizes and figures, and with such other properties, and in such proportions to space, as most conduced for the end for which He formed them ; and that these primitive particles, being solids, are incomparably harder than any porous body, compounded of them ; even so very hard as not to wear or break in pieces ; no ordinary power being able to divide what God himself made one, in the first creation.”

Newton, you will notice, considered all matter as being made up of small solid particles, so “ hard ” as to be unbreakable. But this idea of the nature of matter did not really originate with him, for certain Greek philosophers, many centuries before him, had thought along similar lines. The Greeks were interested in discovering whether matter could be divided and divided indefinitely into smaller and smaller fragments. Some were of the opinion that this could be done. Others held that such a subdivision could not go on indefinitely, but that ultimately we would come to a point beyond which our particle could no longer be broken up. This is Newton’s “ impenetrable particle ” ; and to this particle the Greeks gave the name *atomos* or **atom**, meaning “ which cannot be cut in two.”

48. **The atomic theory.** What the Greeks said and thought about the atom was based upon pure speculation. The Greeks were fond of speculation, but they did little experimental work.

However, much, much later, well within our own time, this Greek idea of "atoms" was revived by John Dalton, an English school-teacher, to explain a number of chemical laws. One of these laws, the *law of definite proportions*, states that elements combine in definite proportions by weight to form compounds. Salt, for example, consists of sodium and chlorine, and when we analyze it quantitatively, we always find the sodium and chlorine present in

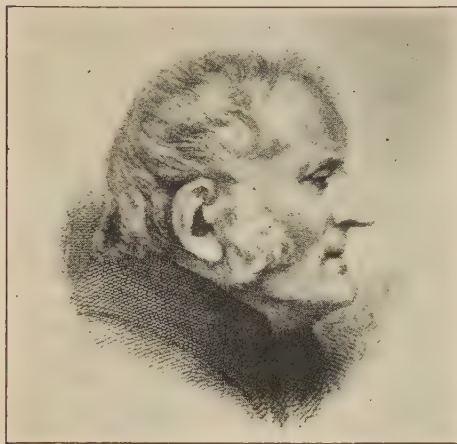


FIG. 54. Dalton.

the proportions of 23 parts of sodium to $35\frac{1}{2}$ parts of chlorine. It does not in the least matter what the source of the salt is: whether we get it from sea water or from salt beds or from any other source. What does matter is that the salt should be absolutely pure salt; and when we analyze this pure salt, we always find the sodium is to the chlorine as 23

parts by weight is to $35\frac{1}{2}$ parts by weight. And the same is true of every one of the 200,000 compounds or so which have so far been prepared: each one of them consists of two or more elements in definite proportions by weight.

49. These were the actual facts in the case — facts which had been gathered by many men in many countries over many years. But man, who is always inquisitive, wanted to know the "why" of it all: why do elements combine in definite proportions by weight? It is Dalton's glory to have supplied us with a satisfactory answer to this question, and to several related questions. He said, let us assume that iron, copper, gold, lead, and the various other elements are made up of very small particles which we, like

the Greeks, shall call atoms. Let us further assume that all the atoms in copper are quite the same, and all the atoms in gold are quite the same, but that the atoms of copper are different from the atoms of gold. We will assume, said Dalton, that one difference between the atoms of different elements is that of weight: the atom of gold does not weigh the same as that of copper, and the atom of copper does not weigh the same as lead, and so on.

50. Now let us turn back to our common salt, which is composed of the elements sodium and chlorine. Dalton assumes that the sodium atom does not weigh the same as the chlorine atom.

Let us assume that the weight of the sodium atom is to that of the chlorine atom as $23 : 35\frac{1}{2}$; in other words, that the chlorine atom is about $1\frac{1}{2}$ times as heavy as the sodium atom. Since the smallest possible part of an element, according to Dalton, is the atom, the smallest part of sodium that can combine with chlorine is one atom, and the smallest part of chlorine that can combine with sodium is one atom. But, according to our hypothesis, one atom of sodium is 23 (parts by weight). Which means that when one atom of sodium combines with one atom of chlorine, the union must be in the proportion of $23 : 35\frac{1}{2}$. And it matters not at all whether we deal with one atom or a billion billion atoms: a billion billion atoms of sodium when combining with a billion billion atoms of chlorine will still continue to combine in the proportions of $23 : 35\frac{1}{2}$.

We have here given a simplified view of Dalton's *atomic theory*, a theory which to this day remains as the foundation stone of chemical science. We believe in atoms just as firmly as the Greeks and Dalton did, but with far more justification, because actual experiments which have been performed since Dalton's day quite

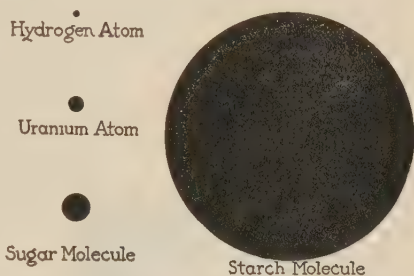
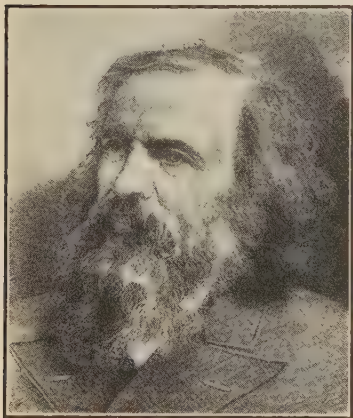


FIG. 55. Sizes of atoms.

definitely prove the reality of atoms. Atoms are as real as you, the reader, and we, the writers. We still believe that atoms are the smallest particles of matter which take part in a chemical change, but we no longer believe that the atoms are the smallest particles possible; because to-day we are familiar with electrons, which are much smaller than atoms. And yet even the atom is small enough: its average diameter is one three-hundred millionth of an inch; and one can pack 77 billion billion atoms of helium in one cubic inch of the gas!

51. The periodic classification of the elements. The next development of importance deals with the classification of the



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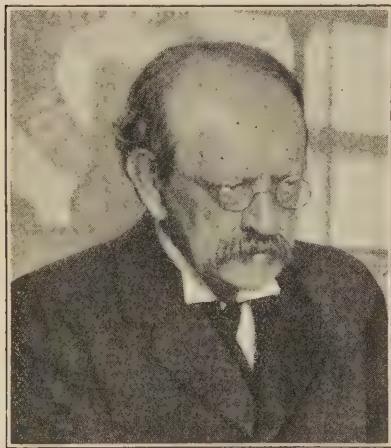
FIG. 56. Mendeléeff.

elements. If we classify the elements according to their density and call the lightest atom, hydrogen, 1, then the atom of sodium becomes 23, that of chlorine $35\frac{1}{2}$, that of oxygen 16, that of nitrogen 14, and so on. The numbers 1, 23, $35\frac{1}{2}$, and 16 represent the *atomic weights* of these elements. Using these atomic weights, Mendeléeff, a Russian chemist, developed his periodic classification of the elements in such a way as to bring out family resemblances.

If you will examine the periodic table (p. 51), you will notice that, except for hydrogen (which does not find a convenient place in the table), he lists the elements in the ascending order of atomic weights, starting with helium 4, lithium 6.94, and so on. When he reaches the eighth element, fluorine, he begins his second line, so that neon falls under helium, sodium under lithium, and so on. Now by such an arrangement — which becomes somewhat more complex as we ascend the table — we discover that the elements in the ver-

PERIODIC TABLE SHOWING ATOMIC NUMBERS AND VALENCES										
Period	Group O	Group I	Group II	Group III	Group IV	Group V	Group VI	Group VII	Group VIII	
O		1 H (1) 1.008								
I	2 He (0) 4.00	3 Li (1) 6.940	4 Be (2) 9.02	5 B (3) 10.82	6 C (4) 12.000	7 N (5) 14.008	8 O (6) 16.000	9 F (7) 19.00		
II	10 Ne (0) 20.2	11 Na (1) 22.997	12 Mg (2) 24.32	13 Al (3) 26.97	14 Si (4) 28.06	15 P (5) 31.027	16 S (6) 32.064	17 Cl (7) 35.457		
III	18 Ar (0) 39.91	19 K (1) 39.096	20 Ca (2) 40.07	21 Sc 45.10	22 Ti 48.1	23 V 50.96	24 Cr 52.01	25 Mn 54.93	26 Fe 55.84	27 Co 58.94
		29 Cu 63.57	30 Zn 65.38	31 Ga 69.72	32 Ge 72.60	33 As (5) 74.96	34 Se (6) 79.2	35 Br (7) 79.916	40 Ni 58.69	
IV	36 Kr (0) 82.9	37 Rb (1) 85.44	38 Sr (2) 87.63	39 Y 88.9	40 Zr 91.	41 Nb 93.1	42 Mo 96.0	43 Tc ?	44 Ru 101.7	45 Rh 102.91
		47 Ag 107.880	48 Cd 112.41	49 In 114.8	50 Sn 118.70	51 Sb (5) 121.77	52 Te (6) 127.5	53 I (7) 126.932		
V	54 Xe (0) 130.2	55 Cs (1) 132.81	56 Ba (2) 137.37	57-71 Rare-Earths	72 Hf 180.8	73 Ta 181.5	74 W 184.0	75 Re ?	76 Os 190.8	77 Ir 193.1
		79 Au 197.2	80 Hg 200.61	81 Tl* 204.39	82 Pb* 207.20	83 Bi* 209.00	84 Po* 210.0	85 —		
VI	86 Rn (0) 222.	87 —	88 Ra*(2) 225.95	89 Ac* 227	90 Th* 232.15	91 Pa* 230?	92 U* 238.17			

tical columns (reading from top to bottom) show striking family relationships. For example, helium, neon, argon, krypton, and xenon show many similar properties; so do lithium, sodium, potassium, and rubidium. But since these and other elements show family relationships, it is conceivable that they all have some common origin; and that therefore not the atom, but something perhaps more fundamental than the atom, is the parent of them all.



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FIG. 57. J. J. Thomson, founder of the electron theory.

52. The electron. Something far more fundamental than the atom was discovered by J. J. Thomson, in the shape of the **electron** (from the Greek word meaning "amber," which gives rise to "negative" electricity when rubbed). We cannot, at this stage, go into the details of this discovery, since the experimental work is of a somewhat complex nature; but enough has now

been done so that we can say with assurance that the electron is a negative particle of electricity; that it is so light that 1700 of them weigh no more than an atom of hydrogen; and that the *same* electron is found in *different* elements, and therefore the electron is more fundamental than the atom, which differs with different elements.

53. The proton. Since the electron represents a negative particle of electricity, one may well ask, where is the positive particle? When the gas from a tube is pumped out and a current then passed through it, electrons shoot out from the cathode. This, in principle, is the kind of apparatus which Thomson used to study the properties of the electron. The cathode representing

the negative pole *repels* the electrons; but at the same time it ought to attract the positive particles of electricity. It actually does do so. By using a perforated cathode, it was shown that cones of violet light projected from the holes behind the cathode. When analyzed, it became apparent that these "rays" could be deflected by magnetic and electrical forces, just as the electrons could, but always in a direction *opposite to that of the electrons*. Obviously, we are here dealing with positive particles of electricity; and these have since been given the name of **protons**. A proton was shown to weigh as much as the hydrogen atom; which is to say, that the proton weighs much more than the electron.

54. The "nucleus" of the atom. The discovery of radium exhibited an element which was liberating enormous quantities of heat, and it was assumed that this was due to the breaking up of the atoms of radium. As in none of the chemical reactions with which we are familiar do we get so much heat evolved, we assume that in chemical reactions the atoms are never "broken," and that only their surfaces, as it were, are affected. A chemical reaction, in fact, is a reaction between atoms, but the nucleus remains intact. This difference between chemical and radioactive changes has led to the view that the atom consists of an inner part or "nucleus," which is never disturbed in any chemical reaction, but which is affected in a radioactive change; and an outer sphere or "planetary circle," which participates actively in chemical reactions.

This view has been much strengthened by experiments due to



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FIG. 58. Ernest Rutherford, to whom we owe our conception of the "nucleus" of the atom.

Rutherford. Using *alpha* particles obtained from radium, he has shown that when these are shot through atoms, they usually continue along a straight line, indicating that there is nothing of a material nature to prevent their passage. But occasionally there is a sudden change in direction, and then we know that these rays have struck some material. This material is the inner part, the "nucleus" or the sun of the atom.

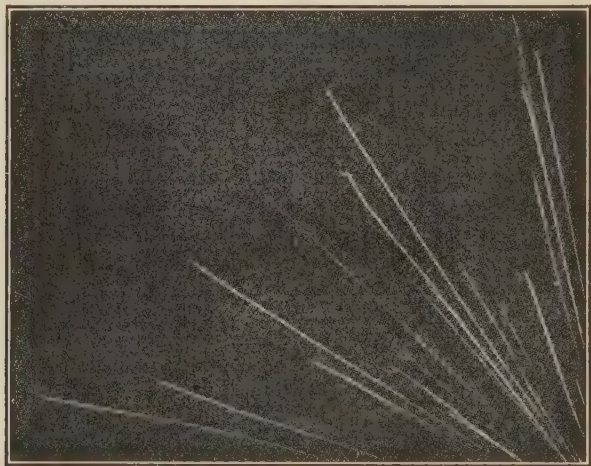


FIG. 59. The tracks of alpha particles.

From the nucleus of radium both *alpha* and *beta* rays are emitted. The alpha rays correspond to the protons and the beta rays to the electrons; which means that both protons and electrons are present in the nucleus. Since, however, the atom as a whole is electrically neutral, and there are electrons (or negative particles of electricity) surrounding the nucleus, we assume that in the nucleus there are more protons than electrons; and the effect of the excess protons in the nucleus is exactly neutralized by electrons surrounding the nucleus.

Since the atom represents a particle of matter, we may say that matter is made up of positive and negative electricity — of protons and electrons.

55. The evolution of the elements. Rutherford has been able to show that at least some of the heavier elements have been built up from some of the lighter ones. For example, by violently bombarding nitrogen atoms with alpha rays, he has obtained hydrogen and helium atoms. This reminds one of the alchemists who for centuries tried to transmute the "base" elements into gold; changing, in other words, one element into another. But what Rutherford had in mind was not to manufacture gold, but to strengthen the theory that most of the elements have probably been built up from the two lightest elements we have, namely, hydrogen and helium.

56. Chemical and radioactive changes. One of the most striking observations made in the study of radium was that nothing, apparently, influenced the rate with which the rays were evolved. Putting radium into an atmosphere of liquid air did not decrease the rate, nor by submitting it to the temperature of the electric furnace was the rate accelerated. Here was something quite different from any chemical change which had ever been studied. A chemical reaction is usually influenced by temperature changes; its rate is usually decreased when the temperature is lowered and increased when the temperature is raised. Obviously, chemical and radioactive changes are fundamentally different. How can we explain this difference?

In radioactive changes the nucleus is affected; in chemical changes only the planetary electrons surrounding the nucleus are affected. In the former, the very heart of the atom, so to speak, is disturbed; in the latter, the nucleus is never touched, but the planetary circle alone is involved. When an atom of sodium combines with an atom of chlorine, their nuclei remain intact; when radium decomposes, its nucleus is disrupted.

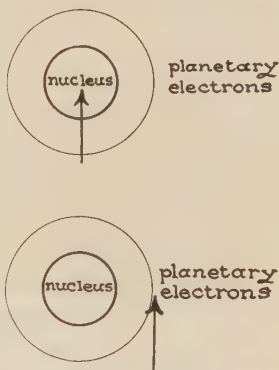


FIG. 60. The upper figure represents a radioactive change, and the lower figure a chemical change.

57. Isotopes. Picture two atoms (1) and (2), with the configurations as shown. (1) is an atom whose nucleus contains 6 protons and 3 electrons; the planetary circle, in B , will, therefore, have to contain 3 electrons so that the atom may remain electrically neutral. Similarly, in (2), since the nucleus contains 7 protons and 4 electrons, the planetary circle, B' , will also contain 3 electrons. Now the *mass* of the atom is centered almost wholly in the

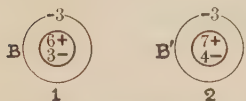


FIG. 61. An illustration of isotopes.

nucleus, since the protons are much heavier than the electrons; in fact, the electrons are almost negligible in weight. If this be so, (2) will weigh more than (1). On the other hand, the chemical properties of atoms depend upon the number (and arrangement) of the electrons in the planetary circles B and B' . Since in our example the number is the same, the chemical properties of these elements would be the same. Here we have an example of **isotopes**: elements with the same chemical properties but with different atomic weights.

Professor Richards, of Harvard, has shown experimentally that lead from radioactive sources has a lower atomic weight than that obtained from ordinary lead ores, though in their chemical behavior the two samples could not be distinguished. Dr. Aston, of Cambridge University, has gone a step further. He has proved that many of our common elements are *mixtures* of isotopes. For instance, ordinary chlorine gas is a mixture of two kinds of chlorine, both having the same chemical properties, and differing only in their atomic weight.



Wide World

FIG. 62. Theodore W. Richards.

ELEMENT	ATOMIC NUMBER	ATOMIC WEIGHT	MINIMUM NUMBER OF ISOTOPES	MASSES OF ISOTOPES IN ORDER OF INTENSITY
Hydrogen	1	1.008	1	1.008
Helium	2	4.	1	4
Lithium	3	6.94	2	7, 6
Beryllium	4	9.02	1	9
Boron	5	10.2	2	11, 10
Carbon	6	12.00	1	12
Nitrogen	7	14.01	1	14
Oxygen	8	16.00	1	16
Fluorine	9	19.00	1	19
Neon	10	20.20	2	20, 22
Sodium	11	23.00	1	23
Magnesium	12	24.32	3	24, 25, 26
Aluminum	13	26.97	1	27
Silicon	14	28.6	2	28, 29
Phosphorus	15	31.03	1	31
Sulphur	16	32.04	1	32
Chlorine	17	35.46	2	35, 37
Argon	18	39.91	2	40, 36
Potassium	19	39.10	2	39, 41
Calcium	20	40.07	2	40, 44
Iron	26	55.84	(1)	56
Nickel	28	58.69	2	58, 60
Zinc	30	65.38	4	64, 66, 68, 70
Arsenic	33	74.96	1	75
Selenium	34	79.2	6	80, 78, 76, 82, 77, 74
Bromine	35	79.92	2	79, 81
Krypton	36	82.9	6	84, 86, 82, 83, 80, 78
Rubidium	37	85.44	2	85, 87
Tin	50	118.7	7	120, 118, 116, 124, 119, 117, 122
Antimony	51	121.77	2	121, 123
Iodine	53	126.93	1	127
Xenon	54	130.2	7	129, 132, 131, 134
Caesium	55	132.81	1	133
Mercury	80	200.61	(6)	(197-200), 202, 204

(Numbers in parentheses are provisional.)

58. Atomic number. In Figure 61 we have an excess of 3 protons in the nucleus. The atomic number, which represents the excess of protons over electrons in the nucleus, is therefore 3 in this case. If the nucleus had 9 protons and 5 electrons, the atomic number would be 4. By photographing the X-ray spectra

of different elements, Moseley showed how the atomic numbers of the elements could be calculated. If the atomic number of hydrogen be called 1, that of helium becomes 2, lithium 3, etc. (4, 5, 6, 7, 8, 9, etc.), until we reach the heaviest element, uranium, which has an atomic number of 92. (See the table on page 51.)

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QUESTIONS

1. Trace the origin of the word "atom."
2. State the law of definite proportions, beginning your sentence with the word "when." Illustrate.
3. Give the important features of Dalton's atomic theory.
4. What is the principle underlying the periodic classification of the elements?
5. Give some essential properties of the electron.
6. How was the proton discovered?
7. What is the "nucleus" of the atom?
8. What is the evidence for the statement that one element can be transformed into another?
9. Contrast chemical with radioactive changes.
10. Explain with the help of diagrams what we mean by isotopes.
11. Define and illustrate, using diagrams, the atomic number.

OPTIONAL QUESTIONS

12. In the light of present knowledge, criticize Dalton's atomic theory.
13. How was the problem of determining the size of the atom approached?
14. Why were isotopes unknown before the development of radioactivity?

CHAPTER VIII

CHEMICAL CHANGE

59. We are now in a position to study a little more closely the processes involved in chemical change. For our purpose we shall use as an example the formation of sodium chloride, or common table salt.

60. Chemical action. Sodium chloride consists of two elements, sodium and chlorine, in chemical combination. Sodium itself is a highly poisonous and extremely reactive metal; chlorine is an equally poisonous, greenish yellow gas. Yet these two can be

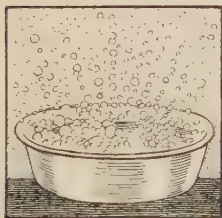


FIG. 63. The action of sodium on water.



FIG. 64. Chlorine gas attack.



FIG. 65. Using sodium chloride as a food.

made to combine to form a product which is not only not poisonous, but is actually an indispensable food. What has happened? How can we picture this change — a change which represents an excellent example of “chemical action”? This we can do by making use of what we have learned in the last chapter. But we must first dwell upon one or two other discoveries.

61. Argon. In 1894, Lord Rayleigh, as a result of many experiments, came to the conclusion that the density of nitrogen obtained from the atmosphere was always slightly, but consistently higher than the density of nitrogen obtained from nitrogen compounds.



Keystone View Co.

FIG. 66. Sir William Ramsay, one of the discoverers of argon and other rare gases of the air.

In conjunction with Sir William Ramsay, he was eventually able to show that this difference was due to an impurity present in the nitrogen obtained from the atmosphere. This "impurity" was separated and isolated. It turned out to be a hitherto undiscovered gas in the air, and the name *argon*, meaning inactive, was given to it because it could not be made to combine with any other element.

62. Stable elements. Soon after argon was discovered, Ramsay discovered other rare gases in the air — helium, neon, xenon, krypton, and radon. They all proved to be inactive chemically: they could not be made to combine with other elements.

Dr. Langmuir, an American scientist who had thought much about the nature of chemical action, turned his attention to these stable elements and advanced the theory that these elements were stable because their atoms had stable configurations. They had, in other words, no tendency to rearrange themselves.

63. An example. Let us, at this point, recall what we learned in the last chapter in reference to the structure of the atom.

In the two atoms, *A* and *B* (Fig. 68), there are an excess of 11 protons in the nucleus of *A* and an excess of 17 in the nucleus of *B*. Since the atoms are



Underwood and Underwood

FIG. 67. Dr. Irving Langmuir.

electrically neutral, we have 11 electrons (-11) in the outer circle of *A* and 17 electrons (-17) in the outer circle of *B*. The chemical properties of the element depend upon the number (and distribution) of the electrons in the outer planetary circle. This means that in a study of chemical action we must know something about these electrons in the planetary circle.



FIG. 68.

64. The X-ray tube. As to the number of electrons in the planetary circle, we have already seen that they must equal the number of excess protons over electrons in the nucleus; they must, in other words, be equal to the *atomic number*, the measurement of which we owe to Moseley, a young English scientist who was killed in the late war.

To determine atomic numbers requires an X-ray tube. In its simplest form this tube, shaped as shown in the diagram, has its

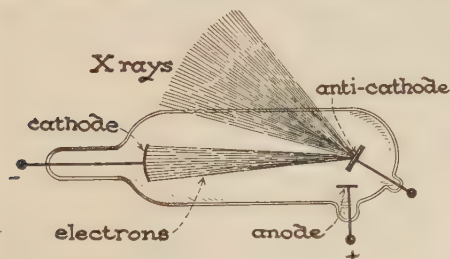


FIG. 69. X-ray tube.

air pumped out and a high voltage current sent through. The electrons which come from the cathode are directed towards a metallic plate, and this impact results in the production of X-rays. These X-rays are considered to

be vibrations in the ether, similar to light and wireless waves; but the rate of vibration of these X-rays is much more rapid than either that of light or wireless waves.

65. Atomic numbers. The rate of vibration of these X-rays varies somewhat, depending upon the metallic plate used. In fact, as we substitute different elements for the metallic plate, we get different rates of vibration. Moseley passed these X-rays through a salt crystal and examined the spectra obtained - a process somewhat similar to passing light through a prism and examining the resulting spectrum. He was able to show that a simple rela-

tionship existed between the wave length and the density of the element; and from this he was able to obtain its *atomic number*, which represented the excess of protons over electrons in the nucleus, as follows:

ELEMENT	ATOMIC NUMBER	ATOMIC WEIGHT
hydrogen	1	1
helium	2	4
lithium	3	7
beryllium	4	9
boron	5	11
carbon	6	12
nitrogen	7	14
oxygen	8	16
etc.	etc.	etc.

The atomic number is, as a rule, one half of the atomic weight (which represents how many times the atom of the element is as heavy as the atom of hydrogen).



FIG. 70. Lewis.

For example, carbon has an atomic weight of 12 and an atomic number of 6; oxygen has an atomic weight of 16 and an atomic number of 8.

66. Langmuir and Lewis. If we know the atomic number of an element, we also know the number of electrons in the outer shell; for, as we have seen, the two must be equal. But how are the electrons in the outer shell distributed? One theory, favored by chemists, has been offered by Langmuir (p. 60) and Lewis (a professor of chemistry at the University of California). They begin with the general proposition that the inert gases of the air (argon, helium, etc.) are inert or inactive because they have stable chemical configurations. Helium, the lightest of

these inert gases, has an atomic number of 2. It may be represented as in Figure 71. It has two planetary electrons, or its outer shell consists of 2 electrons. Since helium is inactive, an outer shell of 2 electrons placed nearest to the nucleus is considered by Langmuir and Lewis to be a stable configuration.



FIG. 71. Helium.

Neon, the next inert gas, has an atomic number of 10. The neon atom may be represented as in Figure 72. But



FIG. 72. Neon.

we have just said that the first shell must contain 2 electrons. On this basis, neon would have two shells, one containing 2 electrons and another containing 8 electrons, as in Figure 73.

Argon, with an atomic number of 18, is the third inert element. Since the first shell must contain 2 electrons and the second 8, we must here have a third shell of 8 electrons to give us a total of 18.

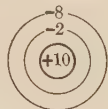


FIG. 73. Neon.

The other inert gases have shells divided as follows: krypton, 2, 8, 8, 18 (atomic number 36); xenon, 2, 8, 8, 18, 18 (atomic number 54); and radon, 2, 8, 8, 18, 18, 32 (atomic number 86).

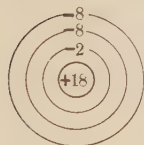


FIG. 74. Argon.

67. Sodium chloride.

We can now return to the question, what happens when sodium and chlorine combine to form sodium chloride? What happens, in other words, in a typical chemical change?

The atomic number of sodium is 11. Its atom may be represented as in Figure 75. Its first shell has 2 electrons, like helium; its second, 8 electrons, like neon; but its third shell has but one electron, and it would need 7 more to make it stable, like argon. Or, it might become stable by losing this one electron and assume the configuration of neon. As a matter of fact, experience shows that in a chemical reaction, it is easier for sodium to lose one electron than to gain 7 electrons.



FIG. 75. Sodium.

Turning our attention to chlorine, with an atomic number of 17, we may represent it as in Figure 76. Here again, for the atom to become stable, one of two things must happen: it must either gain one electron or lose 7 electrons. Experience shows that it is easier for chlorine to gain one electron than to lose 7 electrons.



FIG. 76. Chlorine.

68. Formation of sodium chloride. The diagram explains what happens. Sodium loses one electron and chlorine gains it, thereby creating stability. The

sodium is an active element because it has an electron which it wants to lose, in order to create a stable system; the chlorine is an active element because it needs one more electron in order to create a stable system. What is more natural, than when sodium and chlorine are brought together, that the former should transfer an electron to the latter?

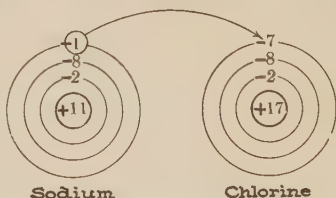


FIG. 77. Chemical combination of sodium and chlorine.

The chemical union, then, of sodium and chlorine consists in the transfer of an electron from the former to the latter.

Notice that in the molecule of sodium chloride we no longer deal either with a sodium atom or a chlorine atom:

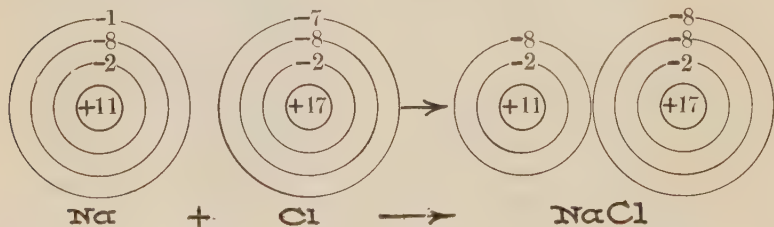


FIG. 78. Formation of the sodium chloride molecule.

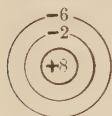
hence we cannot expect to have the properties of sodium chloride resemble the properties of sodium or chlorine,

The chemical union of sodium and chlorine consists, then, in a fitting together of the single outer electron of the sodium atom and the outer shell of seven electrons of the chlorine atom, resulting in a stable shell of eight electrons. Here we have a **molecule** of salt. *It does not contain a sodium atom; it does not contain a chlorine atom; but consists of a new set of shells and therefore has new properties.* This particle, the *molecule*, is the smallest bit of salt that can exist and still retain the properties of salt. Take away even one electron and it is no longer salt.

69. Water. Water consists of two atoms of hydrogen and one atom of oxygen. Its formula is H_2O . It can be made by causing hydrogen and oxygen to combine together. But why do these two elements combine? If we apply similar reasoning here to what we did under sodium chloride, we shall have no difficulty in finding a reason.

FIG. 79.
Hydrogen.

Hydrogen has an atomic number of 1. It may be represented as in Figure 79. It needs another electron in its outer shell to become stable; or it may lose this outer electron in the course of a chemical reaction.

FIG. 80.
Oxygen.

Oxygen has an atomic number of 8. It may be represented as in Figure 80, from which we see that oxygen needs 2 more electrons to create a stable configuration. One hydrogen atom can contribute but one electron; therefore two hydrogen atoms would be needed to supply the two electrons which the oxygen atom lacks. Therefore H_2O .

70. Valence. The number of electrons in the outermost circle gained or lost indicates the valence of the element. Elements which lose electrons are positive and those which gain electrons are negative. For instance, in our pre-

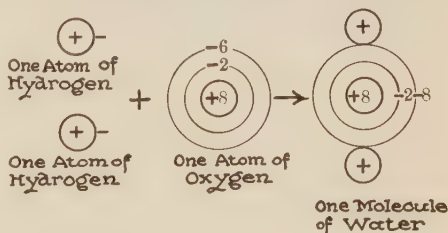


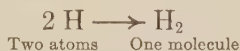
FIG. 81. Formation of water.

vious examples, the valence of hydrogen is + 1, oxygen - 2, chlorine - 1, sodium + 1.

71. Metal and non-metal. Usually, the metals have relatively few electrons in the outermost ring, whereas the non-metals have relatively many. Thus sodium, a metal, has one electron, but chlorine, a non-metal, has 7.

72. Chemical activity. The fewer the number of electrons in the outermost ring, the greater the activity of the element. Sodium with 1 electron is more active than calcium with 2 electrons, and the latter, in turn, is more active than boron with 3 electrons. On the other hand, where there are 7 electrons in the outermost ring, the activity resembles that of sodium with 1 electron; where there are 6 electrons, the activity resembles that of calcium, with 2 electrons; an element with 5 electrons resembles in activity an element with 3 electrons.

73. Atomic hydrogen. With its tendency to discard an electron and become H^+ , or add an electron to form a more stable system, the atom of hydrogen is an extremely active substance. When liberated as a gas (from the action of zinc on sulphuric acid, or from the action of sodium on water), we have reasons for believing that no sooner are the atoms of hydrogen released than they combine in twos to form molecules:



the molecule being much stabler than the atom. That this theory rests on a sound foundation is strengthened by the recent experiments of Langmuir in which he has been able to tear apart the hydrogen molecule into its constituent atoms:



Hydrogen molecules at very low pressures, he has shown, break up into atoms when brought in contact with a tungsten filament at the high temperature of the incandescent lamp.

Professor Wood, of Johns Hopkins, was also able to obtain atomic hydrogen by passing electric discharges through ordinary

hydrogen at low pressures and removing the atoms formed by vacuum pumps.

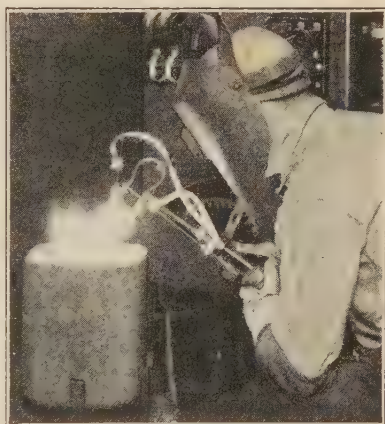
The hydrogen atoms so obtained showed themselves to be far more reactive than hydrogen molecules (which is what ordinary hydrogen as we prepare it in the laboratory consists of). For example, we have seen that when hydrogen is passed over heated copper oxide, copper and water are formed; however, the atomic hydrogen obtained by Langmuir will transform copper oxide into copper at the ordinary temperature.

Such experiments force us to the conclusion that the symbol for the ordinary hydrogen with which we are familiar should be written H_2 and H should be reserved for the more active atomic form. Similarly, it is believed that the symbols for the oxygen and nitrogen with which we are familiar are O_2 and N_2 .

74. Use of atomic hydrogen.

Incidentally, atomic hydrogen has lately been put to a very practical use, thanks to Dr. Langmuir. Hydrogen gas is first passed over an electric arc. The heat of the arc breaks up the hydrogen molecules into hydrogen atoms. These combine again a short distance beyond the arc into molecules of the gas, and in so doing liberate an enormous amount of heat; so that very effective welding can be done by its means. Fluxes are not needed here, and surface oxidation is prevented, since any oxides formed are immediately reduced to metal by the hydrogen.

75. Catalysis. It has recently been shown that metals such as platinum or palladium, when brought in contact with hydrogen



General Elec. Co.

FIG. 82. Atomic hydrogen flames.

gas, convert the hydrogen molecules (H_2) into hydrogen atoms (2 H), which, as has just been pointed out, are chemically very active. A number of very important industrial processes, such as the union of nitrogen and hydrogen to form ammonia and the hydrogenation of fats and the manufacture of sulphuric acid, depend upon the presence of some metal like platinum, iron, or nickel. We speak of these metals as catalytic agents. It is quite probable that, in some cases at least, these catalytic agents are so helpful in chemical reactions because they have the power to convert the relatively inactive molecules of an element into the relatively active atoms of that element. This really gives us a new view of catalysis.

SUPPLEMENTARY READING

(See the previous chapter.)

Guy Bartlett, "The Cathode-Ray Tube," *Journal of Chemical Education*, vol. III (1926), p. 1369.

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B. S. Hopkins, "Ilinium, a New Element," *Science*, vol. LXIII, p. 757 (1926).

Guy Bartlett, "Atomic Hydrogen Flames," *Journal of Chemical Education*, vol. IV (1927), p. 39.

H. S. Taylor, "The Simplest Element of All," *Scientific American*, March (1927), p. 162.

QUESTIONS

1. Contrast the properties of sodium chloride with those of its elements. Explain why the properties of sodium and chlorine have disappeared in the formation of the compound.

2. What do we mean by "stable elements" ("inert" elements), and how did the idea of stable elements help Langmuir and Lewis in formulating their theories?

3. Describe briefly, with the help of a diagram, our present idea regarding the structure of the atom.

4. How does the atomic number differ from the atomic weight?

5. Given that the atomic number of an element is 3, make a diagram to represent its probable structure.

6. Given that the atomic number of hydrogen is 1 and nitrogen 7, show by a diagram what will happen when these two elements combine.

7. Define and illustrate the word "valence."

8. How can we predict from the atomic number whether an element is a metal or non-metal?

9. Given the atomic numbers of elements to be 12, 13, 19, 20, predict, with the help of diagrams, their valency, whether metal or non-metal, and their relative chemical activity.

10. Given the atomic number of calcium as 20 and chlorine 17, with the aid of a diagram show how these two elements combine.

11. Define chemical change using the word "electron."

OPTIONAL QUESTIONS

12. Give a brief account of the discovery of argon.

13. Make drawings representing the atomic structure of the first twenty elements in the periodic system, and in each case indicate the valence, the chemical activity, and whether metal or non-metal.

14. Choose elements from among the first twenty in the periodic table which resemble one another, and with the help of diagrams explain why.

15. Why are the properties of hydrogen chloride different from the properties of hydrogen and chlorine? Show this by means of diagrams.

CHAPTER IX

FORMULAS AND EQUATIONS

76. Those who have studied algebra remember that they had frequent need of symbols and formulas. How familiar is the clause "let $x =$ " ! For a similar reason, in chemistry we have found it necessary to devise a language of our own ; to express, in a simple way, the chemical changes that take place. Most of our symbols correspond to words in a language : C for carbon, O for oxygen, H for hydrogen, etc. One capital letter is used wherever possible, but if the names of two elements begin with the same letter, we cannot use it for both ; so, for instance, C is carbon, Ca, calcium, the "a" being a small written letter, not a capital. Some of the symbols are derived from old Latin or Greek names. Following is a list of the elements whose symbols you will have to use often during your course in chemistry :

VALENCE	NAME	SYMBOL	VALENCE	NAME	SYMBOL
1	Bromine	Br	2	Magnesium	Mg
1	Chlorine	Cl	2	Oxygen	O
1	Hydrogen	H	2	Sulphur	S
1	Iodine	I	2	Zinc	Zn
1	Potassium	K (kalium)	2 or 3	Iron	Fe (ferrum)
1	Silver	Ag (argentum)	2 or 4	Manganese	Mn
1	Sodium	Na (natrium)	3	Aluminum	Al
2	Barium	Ba	3 or 5	Nitrogen	N
2	Calcium	Ca	4	Carbon	C
2	Copper	Cu (cuprum)	5	Phosphorus	P

77. Formulas. When we discover a new substance, one of the things we want to know about it is : what is it composed of, and how much ? Instead of saying : this substance is composed of

two atomic weights of hydrogen and one of oxygen, we write, H_2O , the *formula*, which conveys all that information in a smaller space. In the last century or so, during which chemistry, as a science, has been developing, we have worked out countless formulas, although we do not, by any means, know the formula of every substance.

78. Valence is the factor which determines how many atoms of each element goes into the molecule. You will recall that an element with one electron to lend has a valence of 1, positive; while an element which borrows one electron, has a valence of 1, negative. For instance, sodium chloride is NaCl , because sodium has a valence of 1 positive, while chlorine has a valence of 1 negative. What ought to be the formula of calcium chloride? Calcium has a valence of 2, chlorine 1. Ca will lend its two electrons to two Cl atoms, since each can only receive one electron. The formula of calcium chloride, then, is CaCl_2 . Al (valence = 3) and O (valence = 2) combine to form aluminum oxide, Al_2O_3 , since 2 Al can lend 6 electrons, which are taken by 3 O.

Now, to make this remembering of formulas a simpler matter, if one knows the valences of the elements or groups, proceed as follows: Write the formula of zinc chloride.



(The Roman numerals above the symbols indicate the valence.)

The valence of zinc ($\text{Zn} = 2$) is made the subnumber of Cl and the valence of Cl (1) becomes the subnumber of Zn. (Notice that the subnumber of Zn, 1, is omitted.)

Aluminum oxide,



Where the valences are the same, of course, no subnumbers are used; therefore calcium oxide is CaO .

79. Groups or radicals. In English words, we frequently find groups of letters which cling together, but which, in themselves,

are not words: "ness," as in goodness, "tion," as in convention, etc. So, also, in chemical compounds, we find SO_4 as in Na_2SO_4 , MgSO_4 , the sulphate radical or group; CO_3 as in CaCO_3 , K_2CO_3 , the carbonate group; NO_3 as in NaNO_3 , KNO_3 , the nitrate group; etc. Not one of these is a substance, but the elements composing the group stay together during many chemical changes; and it is convenient to give it a name and think of it as having a valence.

When we say the OH (hydroxyl) group has a valence of 1, we mean



FIG. 83.
OH group.

this: the oxygen atom has gained an electron from the hydrogen atom, but it can take one more; so the group has a valence of 1. Graphically we can represent it as in Figure 83, which shows the hydrogen bound to the oxygen by the common electron while the oxygen is reaching out for one more electron to complete its outer ring. It is easier, of course, to think of the OH group as having a valence of 1, and likewise the other groups:

GROUP	NAME OF GROUP	VALENCE
OH	Hydroxyl	1
NO_3	Nitrate	1
HCO_3	Bicarbonate	1
SO_4	Sulphate	2
CO_3	Carbonate	2
PO_4	Phosphate	3

In writing formulas with these groups, consider the group as one element. For example, the formula of sodium sulphate is



If more than one group is required, a parenthesis, (), must be used: $\text{Mg}(\text{NO}_3)_2$ is magnesium nitrate; $\text{Ca}(\text{OH})_2$ is calcium hydroxide; but NaOH or ZnSO_4 need no parentheses, since there is only one group (just as in algebra, $a + b$, means the sum of a and b — no parenthesis is needed; twice that sum is $2(a + b)$ —

here we need a parenthesis because we wish to express two times "a" and also two times "b").

It is advisable, at this point, to fill out the table called for in question 5 at the end of the chapter, so as to get practice in formula writing.

80. Names of compounds. In naming a compound we try to tell the names of the elements composing it. "Sodium chloride" tells us that the compound is composed of sodium and chlorine. The ending *ide* is added to tell us that there are no other elements. Potassium sulphide — how many elements? The answer is two: potassium and sulphur. Potassium sulphate, K_2SO_4 , has oxygen besides potassium and sulphur. The ending *ate* means oxygen.

$NaCl$, sodium chloride.

$NaClO_3$, sodium chlorate.

Mg_3N_2 , magnesium nitride.

$Mg(NO_3)_2$, magnesium nitrate.

The ending *ite* means less oxygen than *ate*.

Na_2SO_4 , sodium sulphate.

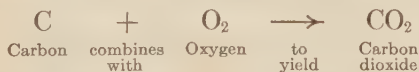
Na_2SO_3 , sodium sulphite.

Where an element has two valences, the lower is called *ous*, the higher, *ic*.

ferrous chloride, $FeCl_2$.

ferric chloride, $FeCl_3$.

81. Equations. In the beginning of this chapter chemical formulas were likened to English words. Then what are sentences? Sentences are equations. That is, whenever you can write a sentence telling the facts of a chemical change, you can also write an equation. The sentence can be translated into chemical language. Suppose we wish to say, in chemical language, that "carbon combines with oxygen to yield carbon dioxide," then we write



Notice, we write formulas for the names of substances.

If we try to write an equation for the decomposition of water, we have $\text{H}_2\text{O} \longrightarrow \text{H}_2 + \text{O}_2$; since H_2 and O_2 are the formulas of the gases, not H and O. But this seems to violate the *Law of the Conservation of Matter*: In any chemical change, the amount of matter produced must be the same as at the start. Have we produced more oxygen? No! we simply have not told the whole story, which is, $2 \text{H}_2\text{O} \longrightarrow 2 \text{H}_2 + \text{O}_2$.

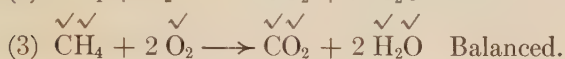
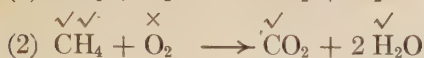
If you recall the electrolysis of water, you will remember that two volumes of hydrogen were produced for each volume of oxygen.

Now we can arrive at the same conclusion by a reasoning process called *balancing the equation*. Starting once more with the unbalanced equation, $\text{H}_2\text{O} \longrightarrow \text{H}_2 + \text{O}_2$, we notice, on the left, two atoms of H (in H_2O); on the right two atoms of H in H_2 . This balances, so far. As to oxygen, on the left, one O; on the right, two O. This does not balance. To get O_2 on the right, we must start with that much on the left. But H_2O has only one O; therefore, we require two H_2O . The equation now reads $2 \text{H}_2\text{O} \longrightarrow \text{H}_2 + \text{O}_2$. We now have 4 H on the left, 2 H on the right. This is balanced by writing 2H_2 ; and we have $2 \text{H}_2\text{O} \longrightarrow 2 \text{H}_2 + \text{O}_2$.

Take another example: the action of zinc on hydrochloric acid:

STEP	DESCRIPTION	ACTUAL OPERATIONS
1.	Write formulas for substances taken and those obtained.	$\text{Zn} + \text{HCl} \longrightarrow \text{H}_2 + \text{ZnCl}_2$
2.	Starting at left end, check number of atoms with those on right, until it fails to check.	$\begin{array}{ccccccc} \checkmark & & \times & & & & \checkmark \\ \text{Zn} & + & \text{HCl} & \longrightarrow & \text{H}_2 & + & \text{ZnCl}_2 \end{array}$
3.	Change amounts, not formulas (that is, never change subnumbers).	
4.	Check again, and continue until balanced.	$\begin{array}{ccccccc} \checkmark & & \checkmark & \checkmark & & \checkmark & \checkmark \checkmark \\ \text{Zn} & + & 2 \text{HCl} & \longrightarrow & \text{H}_2 & + & \text{ZnCl}_2 \end{array}$

Take another example: the complete combustion of CH_4 :



In the examples we have given of balancing equations, notice very carefully that *no one formula, in itself, has been changed*. Water is always H_2O , and hydrochloric acid is always HCl . Balancing an equation very often requires some number in front of the formula; as, $2 \text{H}_2\text{O}$, 2HCl .

There is a temptation to balance equations by changing formulas. In the example $\text{H}_2\text{O} \longrightarrow \text{H}_2 + \text{O}_2$, why not add an extra atom of oxygen to H_2O to balance the equation? Why not? Simply because that would give us hydrogen peroxide, a compound quite different from water.

82. Types of equations. There is an equation for every chemical change. Those that have been covered in the usual laboratory course by this time include

1.	Decomposition (separation of a compound into simpler compounds or elements).	$2 \text{HgO} \longrightarrow 2 \text{Hg} + \text{O}_2$ $2 \text{H}_2\text{O} \longrightarrow 2 \text{H}_2 + \text{O}_2$ $2 \text{KClO}_3 \longrightarrow 2 \text{KCl} + 3 \text{O}_2$
2.	Direct combination of elements.	$\text{H}_2 + \text{Cl}_2 \longrightarrow 2 \text{HCl}$ $2 \text{Cu} + \text{O}_2 \longrightarrow 2 \text{CuO}$
3.	Simple replacement (a free element sets another free, taking its place).	$\text{Zn} + \text{H}_2\text{SO}_4 \longrightarrow \text{ZnSO}_4 + \text{H}_2$ (Zinc takes the place of hydrogen.)
4.	Double replacement, or double decomposition.	$2 \text{NaCl} + \text{H}_2\text{SO}_4 \longrightarrow 2 \text{HCl} + \text{Na}_2\text{SO}_4$ $\text{AgNO}_3 + \text{NaCl} \longrightarrow \text{AgCl} + \text{NaNO}_3$

QUESTIONS

1. According to the electron theory, what is there similar about two atoms that have the same valence?

2. Give names to the following: AgNO_3 ; CaSO_4 ; NaNO_3 ; ZnCl_2 ; $\text{Ca}_3(\text{PO}_4)_2$; FeSO_4 ; Al_2O_3 ; MnO_2 ; CuO ; K_2CO_3 .

3. The valence of Th (thorium) is 4. Write the formula of the oxide.

4. If As_2O_3 is the formula of arsenious oxide, what would be the formula of the chloride?

5. Copy the following tabular form and fill in the formulas

	1 OH	1 Cl	1 NO ₃	2 SO ₄	2 CO ₃	3 PO ₄	2 O	2 S
H					H_2CO_3	H_3PO_4	H_2O	H_2S
Na			NaNO_3	Na_2SO_4	Na_2CO_3	Na_3PO_4	Na_2O	Na_2S
Ca			$\text{Ca(NO}_3)_2$	CaSO_4	CaCO_3	$\text{Ca}_3(\text{PO}_4)_2$	CaO	CaS
Zn			$\text{Zn(NO}_3)_2$	ZnSO_4	ZnCO_3	$\text{Zn}_3(\text{PO}_4)_2$	ZnO	ZnS
Al			$\text{Al}_2(\text{NO}_3)_3$	$\text{Al}_2(\text{SO}_4)_3$	$\text{Al}_2(\text{CO}_3)_3$	AlPO_4	Al_2O_3	Al_2S_3
Fe			$\text{Fe(NO}_3)_3$	$\text{Fe}_2(\text{SO}_4)_3$	$\text{Fe}_2(\text{CO}_3)_3$	$\text{Fe}_3(\text{PO}_4)_2$	Fe_2O_3	Fe_2S_3
K								
Ag								

6. Explain when a parenthesis is used in a formula. What does the parenthesis mean in the formula $\text{Al}_2(\text{SO}_4)_3$? Why is no parenthesis used in the formula Na_2SO_4 ?

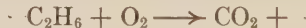
7. Write formulas for mercurous chloride and mercuric chloride; ferrous oxide and ferric oxide.

8. Write equations for

- Heating copper in air.
- Heating carbon in air.
- Heating sulphur in air.
- Heating hydrogen in air.

9. What class of compounds is formed by heating substances in air?

10. Complete and balance the following equation



11. Write the following equations:

- Burning iron.
- Preparation of oxygen from Na_2O_2 .
- Preparation of hydrogen from HCl .
- Action of sodium on water.
- Action of calcium on water.

12. Write equations for

- a.* Reduction of copper oxide.
- b.* Reduction of iron oxide.
- c.* Explosion of hydrogen.

13. Go over your experiments and demonstrations and write an equation for each chemical change you have observed.

OPTIONAL QUESTIONS

14. Write an equation to represent what might be expected to take place when zinc sulphide is roasted (heated in air).

15. Write an equation for the complete combustion of C_2H_2 ; CH_3OH .

16. Give all the information which this equation conveys to you :



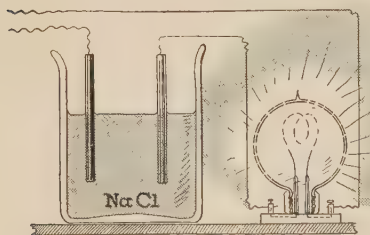
17. Write an equation not given in the text for

- a.* Decomposition.
- b.* Combination.
- c.* Replacement — Simple.

CHAPTER X

CHEMICAL ACTION IN SOLUTION

83. When ordinary dry substances are brought together, very little action takes place; chemical change is very much increased in the liquid state, due to the greater freedom of the molecules;



and since solution is also a liquid, we shall expect, here also, increased reactivity.

84. Kinds of solution.

If we confine our attention to solutions in liquid media, like salt water, sugar solution, and milk, we find we are dealing with different things. In the first place, milk seems to be a kind of solution, easy to distinguish: it is cloudy. It is, in fact, a colloidal solution, a type so numerous and important, that it will be taken up in a separate

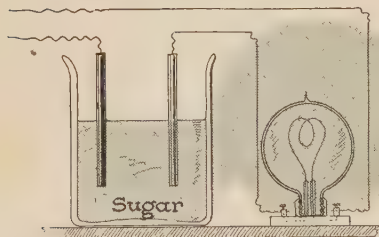


FIG. 84. The top figure illustrates the action of a conductor; the bottom figure that of a non-conductor.

chapter. The sugar and salt solutions look like plain water—the sugar and salt being completely invisible. Upon investigation, however, it is found that the salt water will conduct an electric current whereas the sugar solution will not. We note, then, two

classes of solution easily detected by electrical conductivity, as follows :

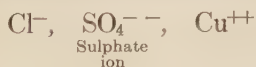
ELECTROLYTES (CONDUCTORS)	NON-ELECTROLYTES (NON-CONDUCTORS)
Acids Bases Salts	Pure water Sugar Alcohol Glycerine

85. Ionization. An explanation of the abnormal behavior of electrolytes was offered, in 1887, by Arrhenius. It had already been shown that the effects of the solute on the solvent (depression of freezing point, increase in boiling point, etc.) were proportional to the numbers of molecules; and since electrolytes had greater effects, Arrhenius concluded there must be more molecules (or particles). Now it was known that during the passage of the electric current, — for instance, through a solution of copper chloride, CuCl_2 , — decomposition takes place, forming a new substance at the anode (+ pole or electrode), and also at the cathode (— pole or electrode). For instance, in the case of CuCl_2 , Cu is deposited at the cathode, and chlorine at the anode. It is also true that many substances dissociate at higher temperatures: $2\text{HI} \longrightarrow \text{H}_2 + \text{I}_2$; $\text{NH}_4\text{Cl} \longrightarrow \text{NH}_3 + \text{HCl}$. Arrhenius therefore concluded that in the solution of CuCl_2 , each molecule was broken up into particles of copper and particles of chlorine called *ions*, “wanderers”; that these ions acted like molecules in their effect on the solvent, and furthermore, carried the electricity through the solution during electrolysis. To sum up :

1. Electrolytes break up or *dissociate* into ions during solution :



2. Ions are electrically charged atoms or groups of atoms :



3. On the passage of the electric current, the $+$ ion moves toward the cathode ($-$ pole); the $-$ ion toward the anode ($+$ pole).

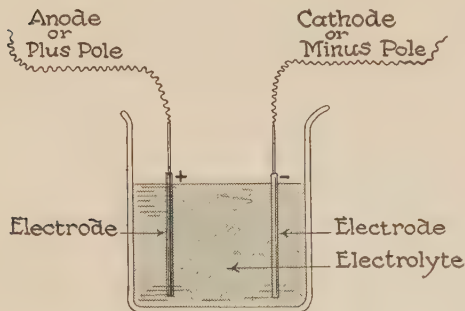


FIG. 85.

These ideas were called by Arrhenius the *theory of electrolytic dissociation*, although it is also referred to as the *Ionization Hypothesis* or *Theory*. It enables us to explain many of the facts of solution which cannot be explained in

other ways, although it still leaves many things in the dark.

86. Ionization and the electron theory. We have already seen that chemical change may be interpreted as a transfer of electrons from the outer shell of one atom to that of another, in order to form stable configurations or shells (par. 67). We can apply that idea to ionization. Let us consider, for example, NaCl, sodium chloride. Solid NaCl consists of ions, each one of which has a stable electron structure, the extra electron of the sodium atom with the seven electrons of the chlorine atom making

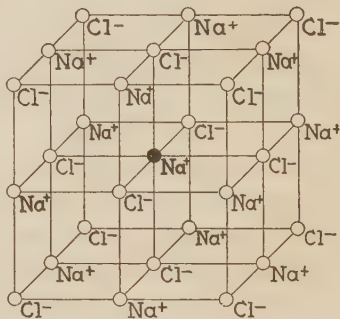
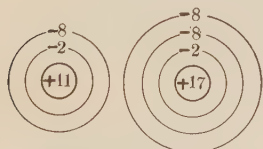


FIG. 86. Solid sodium chloride.



Sodium Chloride

FIG. 87. The sodium chloride molecule.

a complete system of eight electrons, as in Figure 87.

Now, on solution these two parts of the molecule, bombarded by the freely moving water molecules, are separated into the two stable groups as in Figure 88.

It will be noticed that the extra electron of the original neutral atom of Na has gone over to the Cl atom, leaving the Na atom with an excess of one + charge, while the Cl atom now has one electron (− charge) too many. In other words, Na ion is Na^+ , while Cl ion is Cl^- . We have already learned that elements whose outer shell have a stable grouping, like helium, neon, and argon, are chemically inert; that is, there is no tendency to lose or gain electrons. This may explain the chemical inertness of Na^+ (sodium ions) compared to Na atoms. The latter have incomplete outer shells, and are therefore more active than the ions whose outer shell is complete. Sodium atoms act vigorously on water and acids,

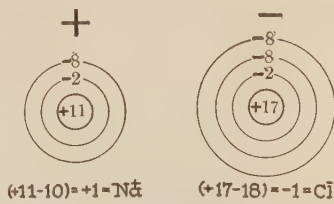


FIG. 88. Sodium ion and chloride ion.

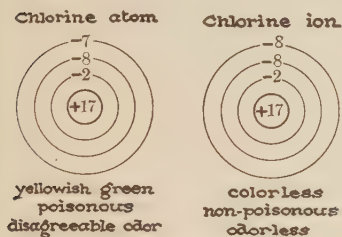
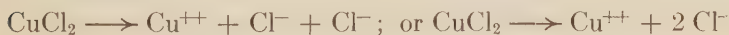


FIG. 89. The chlorine atom and the chlorine ion.

while the ions do not. Similarly, chlorine atoms are active: they combine with metals with evolution of light; the molecules of chlorine have a disagreeable odor and are, in fact, poisonous, none of which properties are exhibited by the ions of chlorine. Of course not; chlorine ions are, for most purposes, a

different substance, having a different number of electrons; just as different from chlorine atoms as gold is from iron.

87. Electrolysis. We have already learned that when an electric current is passed through a solution of an electrolyte, chemical changes take place. We must now study electrolysis in the light of ionization. A simple example is the electrolysis of a solution of copper chloride. On dissolving copper chloride, many of the molecules break up into ions, as follows:



If, now, an electric current is passed through the solution, Cu^{++} is attracted to the cathode (unlike charges attract; like charges repel), while Cl^- is

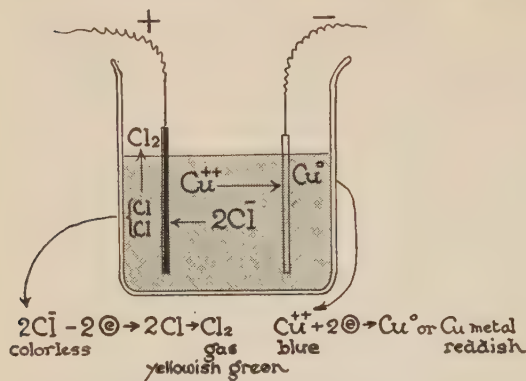


FIG. 90. Electrolysis of a solution of copper chloride.

electrons which convert it to Cu^0 , metallic copper, which clings to the cathode. Copper ions have a blue color and as the electrolysis proceeds, the solution becomes less and less blue, showing the presence of fewer Cu^{++} .

The electrolysis of sodium chloride solution presents a slight variation. Here it will be better to consider the water also as slightly ionized into H^+ and OH^- . We have present, therefore, in a solution of sodium chloride, four kinds of ions, Na^+ , Cl^- , H^+ , and OH^- . On passing the current, Na^+ and H^+ go to the

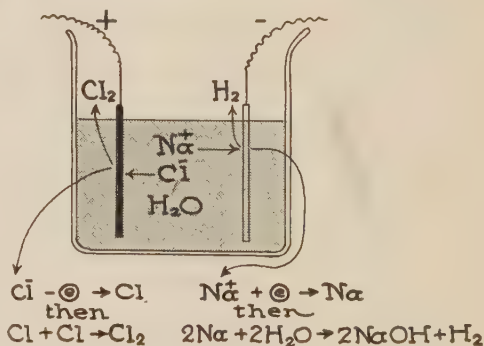


FIG. 91. Electrolysis of a solution of sodium chloride.

cathode, but H^+ , being more easily discharged, forms atoms, then molecules, and bubbles off as hydrogen gas; but no metallic

sodium will be formed so long as any H^+ (in other words water) is present. Others prefer to consider this as a secondary reaction. The sodium ion is discharged, becoming sodium atom, which reacts with water, liberating hydrogen :



At the other pole, OH^- and Cl^- appear, and the Cl^- is more easily discharged. This leaves in solution all the Na^+ and all the OH^- , and when all the H^+ and Cl^- are discharged, we shall have left a solution of NaOH . This is, in fact, the commercial method of making NaOH by the Nelson Process, and, at the same time, chlorine gas, used to make bleaching powder, is obtained as a by-product.

88. Electrolysis of water. In the electrolysis of water, a little sulphuric acid is used (par. 37). Why? The sulphuric acid breaks up into H^+ and SO_4^{--} , the water into H^+ and OH^- . On electrolysis, H^+ is discharged, producing hydrogen gas; while at the anode, OH^- is more easily discharged than SO_4^{--} , forming O and H atoms in equal numbers, which form water and oxygen gas :



leaving in solution the same amount of sulphuric acid as at first. The same thing can be done with sodium hydroxide or sodium sulphate.

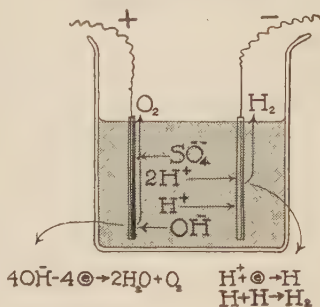
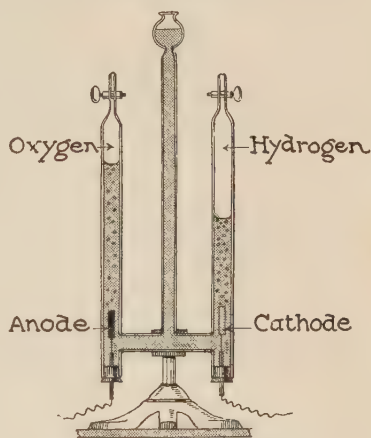


FIG. 92. Electrolysis of water.

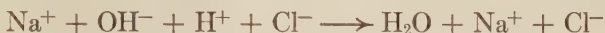
89. Chemical activity of acids and bases. It has already been seen that H^+ tends easily to acquire an electron, while OH^- tends to lose one. This makes these two ions take part frequently in chemical changes. *Substances which in solution have H^+ exhibit acid properties; those which have OH^- show basic properties.* Strong acids contain a high concentration of H^+ , weak acids have less; that is, strong acids are more highly ionized than weak acids, as shown in the table:

ACIDS		BASES	
Nitric	90%	Potassium hydroxide . .	90%
Hydrochloric	90%	Sodium hydroxide . . .	90%
Sulphuric	60%	Barium hydroxide . . .	77%
Oxalic	50%	Ammonium hydroxide . .	1%
Phosphoric	25%	Water	.00001%
Tartaric	10%		
Citric	9%		
Acetic	.1%		
Boracic (boric)	.01%		
Water	.00001%		

90. Strong vs. concentrated. The expressions "strong acid," "strong base" are frequently used when the speaker means concentrated acid, concentrated base. Most people say that sulphuric acid is strong when it contains very much H_2SO_4 , and very little water; for instance, 95% H_2SO_4 . But such an acid is not so strong, that is, not so active, as 80% H_2SO_4 , because it is not so highly ionized. The strength or activity of acid is measured by the per cent ionization or concentration of H^+ , and not by the concentration of H_2SO_4 molecules, since the molecules are inactive. The thick, oily, concentrated sulphuric acid will hardly react with zinc, because it contains relatively few H^+ ; but add a little water, and vigorous action ensues, because now there are more H^+ . To sum up: *the strongest acid is one which has the highest concentration of H^+ . The most concentrated acid has the least amount of water.*

The taste as well as the effect on litmus is due to H^+ for acids, and OH^- for bases. In fact, all the common properties of acids are properties of H^+ , while properties we call basic are due to OH^- .

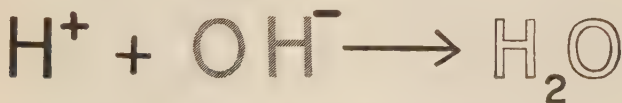
91. Neutralization. The chief property of acids and bases is their tendency to combine with each other. This is due entirely to the great tendency of H^+ to gain and OH^- to lose an electron. For instance,



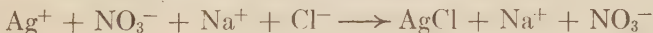
The original solution contained Na^+ and Cl^- , and the final solution contains the same ions. This is true in all cases of neutralization:



and in fact, for each molecule of water formed the same amount of heat is liberated, during any neutralization, irrespective of the acid and base used. This seems to show that the same change is taking place, since the same amount of energy is produced. That is, *neutralization is the process in which H^+ and OH^- unite to form H_2O .*



92. Precipitation. All crystalline materials have a definite solubility; that is, when the solution is saturated, no more will dissolve. Let us consider silver chloride with a solubility, approximately, of .001; that is, saturation is reached when 1 gram of silver chloride is present in 1000 grams of solution. Now suppose we mix solutions of silver nitrate and sodium chloride, both very soluble and highly ionized. A reaction occurs:



We have here large amounts of Ag^+ and Cl^- reacting, which would produce a relatively large amount of $AgCl$; but this is not very soluble. What happens then? Just as soon as there has been

formed one gram of AgCl per 1000 grams of solution, the solution is saturated, and any more that forms cannot remain in solution, and therefore comes out as solid AgCl , which we call a precipitate. And can we predict such a result? At this time we can make a prediction which would be fairly correct. Look up the "Table

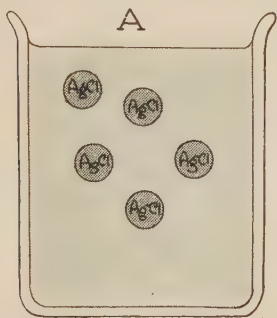


FIG. 93. A saturated solution of silver chloride.

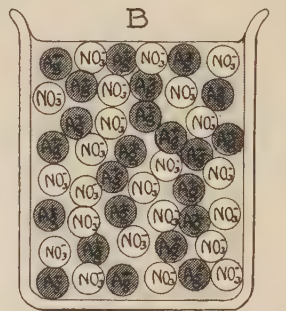


FIG. 94. A solution of silver nitrate.

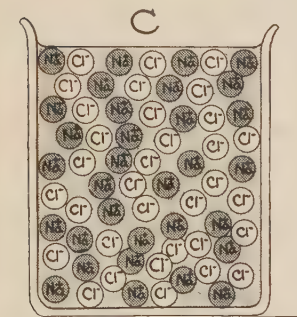


FIG. 95. A solution of sodium chloride.



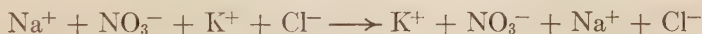
FIG. 96. When B and C are mixed we get a precipitate of AgCl .

of solubilities," in the Appendix, to find out whether one of the possible combinations of ions is an insoluble substance; if so, then that substance would be precipitated. At the same time, if a possible product is a gas or substance which changes to gas easily — in other words, *volatile*, — like carbonic acid, $\text{H}_2\text{CO}_3 \rightarrow \text{H}_2\text{O} + \text{CO}_2$ [carbon dioxide]), then that substance will be formed. Or

if on changing the temperature, or other conditions, one substance can be made insoluble, be it solid or gas, then that one will be formed if the conditions are made right. (For example, HCl is made by raising the temperature.) We must not, however, forget that *no substance is absolutely insoluble*. When a precipitate can be produced, we call that substance insoluble. Such a substance as barium sulphate, considered as one of the most insoluble substances, is still soluble to the extent of one part in 400,000 parts of water. There are, then, three conditions under which a given substance can be formed from a mixture of its ions.

1. *Neutralization.*
2. *Precipitation or insolubility.*
3. *Formation of a gas, or volatility.*

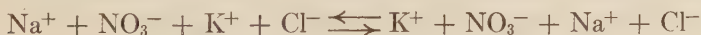
93. Chemical equilibrium and mass action. The three principles just stated lead us naturally to a consideration of the factors controlling all chemical actions, and more especially the velocity of reaction. If the conditions do not belong under 1, 2, or 3 of the last paragraph, then what happens? Apparently nothing will be formed; that is, not enough to produce more than saturation. So whatever is formed remains in solution, and we think nothing has happened. For example,



All four substances are highly ionized and very soluble; therefore, to saturate the solution requires very much material, and no precipitation takes place. But the reaction does proceed to a certain extent; rapidly at first, then more and more slowly, until no further change goes on; that is, *equilibrium* has been established. Exactly the same point may be reached from the other direction:



and so, such reactions are written

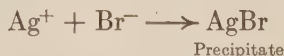


the double arrow indicating equilibrium; that is, the saturation point for any of the four possible substances is not exceeded. We must not, however, forget that even when reactions go to completion, as in the three cases mentioned above, there is still an equilibrium, since the solution remains saturated with respect to the substance formed. We may, then, represent any equation with the double arrow, $\rightleftharpoons$, and it only remains to discover how to force the reaction to right or left and control its velocity.

Let us study the reaction



that is, CuBr_2 , molecular, is brown, whereas Cu^{++} , ion, is blue, as shown in solutions of all copper salts. On the addition of water, the solution becomes more and more blue, and less and less brown, showing increased ionization. If we add sodium bromide to the solution, it becomes brown, showing that the reaction has gone to the left, due to increased concentration of Br^- . Copper sulphate will produce the same effect, due to increased concentration of Cu^{++} . This illustrates a principle: *To force a reaction, increase the concentration of one or more reacting ions.* For instance, in the example given, adding Cu^{++} or Br^- forces the reaction to the left. If now we add to the brown solution a little silver nitrate, it becomes blue. What has happened? The Ag^+ has combined with Br^- according to the equation



and since this is only slightly soluble, it precipitates and removes most of the Br^- . Immediately, to maintain equilibrium, CuBr_2 molecules ionize to form more Cu^{++} and Br^- , and the increased Cu^{++} makes the blue color. This gives us another principle: *To force a reaction, remove one or more of the ions produced.* In this case removing Br^- forced the reaction to the right.

94. Common ion effect. Sodium chloride is very soluble in water but not in hydrochloric acid solution. Why? Consider the equation:



On addition of concentrated hydrochloric acid (better yet, HCl gas) to a saturated solution of sodium chloride, we have increased the Cl^- , forcing the reaction to the left, and making more molecules of NaCl. But the solution is already saturated with NaCl; therefore the excess of molecular NaCl precipitates. This is known as the common ion effect. (In the example given, the common ion is Cl^- .)

95. Hydrolysis. Why does a solution of sodium carbonate, which does not apparently contain OH^- , turn litmus blue? We have already seen that only OH^- produces that effect. Let us study the equation for the formation of sodium carbonate:

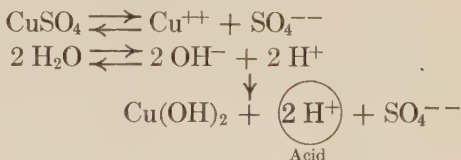


We usually say that water is un-ionized. For all ordinary purposes, this is near enough to the truth. In reality, no substance is absolutely un-ionized. Even water ionizes slightly (.00001%) into H^+ and OH^- , and if we remove water from the solution, we force the reaction to the right, producing more and more Na_2CO_3 , and less and less NaOH and H_2CO_3 . But if we add water, the reaction is forced to the left, Na_2CO_3 breaking up. Now H_2CO_3 is not highly ionized; its saturation point is easily reached and molecules form: $2 \text{H}^+ + \text{CO}_3^{--} \rightleftharpoons \text{H}_2\text{CO}_3$; removing H^+ from solution, and leaving OH^- which now exceeds H^+ ; and since a neutral solution must contain equal concentrations of H^+ and OH^- , this solution is no longer neutral, but basic. Similarly, most salts are really not neutral in solution: CuSO_4 , ZnCl_2 , AlCl_3 , for example, are acid in solution; while solutions of Na_2CO_3 , NaHCO_3 , $\text{Na}_2\text{B}_4\text{O}_7$, turn litmus blue. NaCl, K_2SO_4 and, in general, *salts formed by neutralization of strong bases and strong acids are practically neutral. Salts containing the radicals of a weak acid and strong base make*

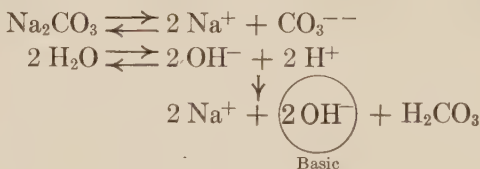
solutions which turn litmus blue. These salts are called **alkalies**. Those containing radicals of a strong acid and weak base produce acid solutions.

Summarizing these ideas:

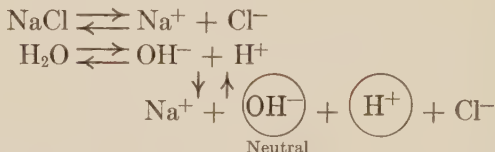
1. *Hydrolysis of a salt which shows an acid reaction.*



2. *Hydrolysis of a salt which shows a basic reaction.*



3. *A salt which is neutral.*



96. Mass action. The principles for the control of reactions are derived from the Law of Mass Action, whose exact statement is difficult for the beginner to understand. The idea, however, is something like this: if an equilibrium is upset by addition or removal of one or more reacting ions (or masses), then further changes will take place in order to reestablish the original equilibrium.

OPTIONAL PARAGRAPH

97. The Law of Mass Action is a mathematical or exact statement. If we confine ourselves to compounds in which valences

are all one, like HCl, NaCl, etc., the law may be expressed as follows:

$$\frac{C_{A^+} \times C_{B^-}}{C_{AB}} = K \text{ (a constant)}$$

which may be translated: in the reaction, $AB \rightleftharpoons A^+ + B^-$
Molec-
ular Ion Ion

the concentration of A^+ multiplied by the concentration of B^- divided by the concentration of molecular AB (un-ionized) is always the same. An example will perhaps make it clear. Let us assume simple whole numbers to represent the concentrations;

our equation might then become $\frac{4 \times 4}{2} = 8$. According to the

law the number 8 is a constant and therefore cannot vary. If we change one factor on the left side, the others must change in such a way as to make the result 8 — always 8. Suppose we change

C_{A^+} from 4 to 8. We now have $\frac{8 \times 4}{2} = 8$, which is not true.

So what happens? C_{B^-} and C_{AB} must change also; for example, like this: $\frac{8 \times 3}{3} = 8$, and this is true.

Let us apply the law to common ion effect to explain why concentrated HCl added to a saturated solution of NaCl forms a precipitate. The mass law for this reaction ($NaCl \rightleftharpoons Na^+ + Cl^-$) is

$$\frac{C_{Na^+} \times C_{Cl^-}}{C_{NaCl}} = K$$

We now add H^+Cl^- , that is, we increase, perhaps double, the C_{Cl^-} and, in order to keep K constant, we must diminish C_{Na^+} ; but this can only be done by causing Na^+ to unite with Cl^- to form NaCl molecules; that is, the denominator C_{NaCl} is increased. The solution, however, is saturated with NaCl; that means, it cannot hold any more, and so the excess must precipitate.

The law of mass action is very useful, because with a table of ionization constants (K), one can predict just how far a reaction will go.

QUESTIONS

1. Explain why an excess of H_2SO_4 is used to precipitate all the Ba from a solution of BaCl_2 (common ion effect).

2. Write an equation to show equilibrium between solid sodium nitrate and its solution.

3. Explain why a mixture of solutions of NH_4Cl and NH_4OH is less basic than a solution of NH_4OH alone.

4. Sodium acetate is much more highly ionized than acetic acid (hydrogen acetate). What might be the effect on the ionization of the acetic acid of adding some solid sodium acetate? How would this affect the strength of the acid?

5. Dilute sulphuric acid conducts the electric current much better than the concentrated acid. What does this show?

6. Write equations to show

a. Solution of NaOH , a strong base.

b. Solution of HNO_3 , a strong acid.

c. Mixing of "*a*" and "*b*."

7. State three properties by which you could identify chlorine gas. Are these also properties of the chloride ion, as shown in a solution of sodium chloride, common salt?

8. Name three ions which are composed of single atoms; two composed of groups of atoms.

9. Draw a diagram to show the electron structure of a potassium ion, and explain why it is inert compared to the potassium atom.

10. Why does FeSO_4 solution turn litmus red?

11. Why does not stearic acid, an insoluble solid, turn litmus red?

12. Dilute H_2SO_4 acts vigorously with Zn. Concentrated H_2SO_4 hardly acts at all. Why?

13. In a mixture of the following substances in solution, name two probable products: HCl , Na_2SO_4 , $\text{Ca}(\text{OH})_2$, Cu , MgCl_2 . In each case tell why you think so. (See par. 92)

14. Why does K_2CO_3 solution turn litmus blue?

15. Why does not Na_2SO_4 solution affect litmus?

16. Write ionic equations for

a. Action of metal on acid.

b. Action of metal on water.

17. What is the difference between strong and concentrated acetic acid?

18. In the preparation of oxygen, why does the reaction go to completion?

OPTIONAL QUESTIONS

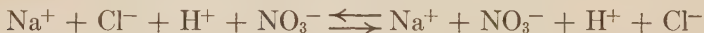
19. How could you identify chloride ions?
20. Name three chemically active metals; two less active ones. What would you expect as to the chemical activity of compounds of these metals?
21. Draw a diagram, and explain the electrolysis of water, by the electron theory.
22. Write an ionic equation for the electrolysis of water.
23. What effect will a solution of mercuric chloride have on litmus? Why?
24. Methyl orange turns red in acid and yellow in basic or neutral solutions. When sodium acetate is added to a solution of acetic acid, the methyl orange turns from red to yellow. Explain.

CHAPTER XI

ACIDS, BASES, AND SALTS

98. In the preceding chapter, we have learned to look upon an acid as a substance which in solution contains free H^+ , and a base as a substance which in solution contains free OH^- . In this chapter it is our intention to study the subject a little more from the practical side.

99. How to make acids. In the preparation of hydrochloric acid (HCl), sodium chloride and sulphuric acid are used, the sodium chloride furnishing the Cl and the sulphuric acid the H. Now if we mix sodium chloride and nitric acid, the following equation would represent the reaction :



that is, equilibrium would soon be established, and while we have some H^+Cl^- in solution, it is of no use to us, since it is mixed with three other substances : $Na^+ + Cl^-$, $H^+ + NO_3^-$, and $Na^+ + NO_3^-$.

What we want is the pure HCl. How are we going to separate it ? In other words, how can we force the reaction to the right ? The Law of Mass Action tells us we must change the conditions until either HCl or $NaNO_3$ is not highly ionized ; then it will separate. We know of no way to accomplish this. Raising the temperature cuts down the dissociation of HCl, but, even more, of HNO_3 . So on heating, what is formed ? HNO_3 ; and we are at a standstill. But when we use sulphuric acid, raising the temperature does not affect its dissociation appreciably, and the HCl is formed. Furthermore, sulphuric acid is cheap and therefore ideal for practical purposes. To make an acid, then, we need two substances :

1. Sulphuric acid (high boiling point).
2. A compound containing the non-metal or group of the required acid :

TO MAKE	WE NEED
HCl (hydrochloric)	$H_2SO_4 + NaCl$ chloride
HNO_3 (nitric)	$H_2SO_4 + NaNO_3$ nitrate
$H_2C_2O_4$ (oxalic)	$H_2SO_4 + PbC_2O_4$ oxalate
$HC_2H_3O_2$ (acetic)	$H_2SO_4 + Ca(C_2H_3O_2)_2$ acetate
$H_2C_4H_4O_6$ (tartaric)	$H_2SO_4 + K_2C_4H_4O_6$ tartrate

100. Properties of acids (that is, of hydrogen ion, H^+).

1. Taste sour.
2. Turn litmus red.
3. Neutralize bases.
4. Give up H_2 (gas) for metal (replacement).

These properties are characteristic only of what may be called "strong acids"; that is, those which ionize to produce more than a certain C_{H^+} (concentration of H ion). Many acids, like stearic, have no sour taste and do not affect litmus, because, to do that, they must have a certain C_{H^+} , and they have not.

SOME COMMON ACIDS	WHERE FOUND
Citric	Lemon, orange
Tartaric	Grape
Malic	Apple
Lactic	Sour milk
Boric	Hot springs in volcanic regions
Acetic	Vinegar

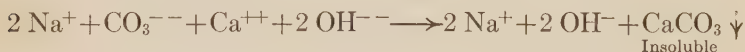
101. Bases. Just as acids depend for their characteristic properties upon H^+ , so bases depend upon OH^- .



FIG. 97. These foods are sour. They contain acids.

102. How to make bases. There is no single method of making bases corresponding to that for most acids. We must get OH^- and that can only be obtained from another base or from water. Sodium hydroxide can be

made by mixing slaked lime, $Ca(OH)_2$, with sodium carbonate, Na_2CO_3 , when the following reaction takes place:



(It is customary to use an arrow sign pointing downward to indicate insolubility of the substance.)

It is the insolubility of the $CaCO_3$ which causes the reaction to go to completion.

But this method has been superseded by the electrolysis of brine:

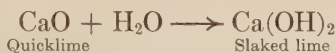


(It is customary to use an arrow sign pointing upward to indicate the volatility of the substance.)

Potassium hydroxide, KOH , is made in the same way as $NaOH$, using potassium compounds in place of the sodium compounds. Of course, these bases can be made by the action of a metal and water:



but this method is impractical, because of the cost of the metal. Calcium hydroxide, $\text{Ca}(\text{OH})_2$, is made very simply from calcium oxide (quicklime) :



BASES	FORMULA	COMMON NAMES
Sodium hydroxide	NaOH	Lye, soda lye, caustic soda
Potassium hydroxide	KOH	Potash, potash lye, caustic potash
Calcium hydroxide	$\text{Ca}(\text{OH})_2$	Slaked lime, milk of lime, limewater
Magnesium hydroxide	$\text{Mg}(\text{OH})_2$	Milk of magnesia
Ammonium hydroxide	NH_4OH	Ammonia water, spirits of Hartshorn

103. Properties of bases (that is, of hydroxyl ion, OH^-).

1. Turn litmus blue.
2. Neutralize acids.
3. Have slippery feeling.
4. Taste soapy (alkaline).

Just as with acids, there are many bases which have not these properties, because the C_{OH^-} is too low; for instance, ferric hydroxide, $\text{Fe}(\text{OH})_3$; zinc hydroxide, $\text{Zn}(\text{OH})_2$; and copper hydroxide, $\text{Cu}(\text{OH})_2$.

104. Uses of bases. The strong bases have the property of dissolving many organic substances, like grease, and so their chief uses are in cleaning. Sodium hydroxide will remove paint and varnish from wood, and calcium hydroxide is used to soften hair preparatory to its removal from hides, before tanning. Neutralization of acids also claims a large portion of base. After sulphuric acid has been used in purifying petroleum, it must be removed by neutralization with a base. "Acid mouth" is treated with magnesium hydroxide (milk of magnesia), and sometimes the acidity of cow's milk is neutralized by calcium hydroxide. Most



FIG. 98. A strong base commonly used in the home.

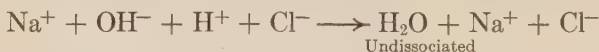
of the sodium hydroxide is used to make soap, the great cleanser, and ammonium hydroxide is used almost exclusively for household



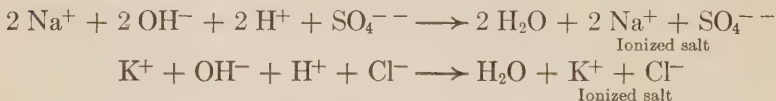
FIG. 99. Basic substances in common use.

cleaning. Great care must be exercised in the choice of a base. Sodium and potassium hydroxides are very powerful bases and caustic in the highest degree, that is they attack the skin and other tissues. Calcium, magnesium and ammonium hydroxides are much milder: ammonium hydroxide, in fact, is volatile and therefore quite safe to handle.

105. Neutralization. This has already been explained as the process in which an acid and base unite to form water, only the H^+ and OH^- taking part in the change. For example:



These reactions go to completion because of the exceedingly slight dissociation of water, so that out of the original four ions in our example, there remain only two: Na^+ and Cl^- . Now this is exactly what we have when we dissolve $NaCl$ in water; in other words, a salt solution. In every case of neutralization, then, we get water and a salt solution. A few other examples are here given:



A salt, then, may be defined as the product, other than water, obtained by neutralization. The same substance may also be obtained by replacing H_2 from an acid by a metal:



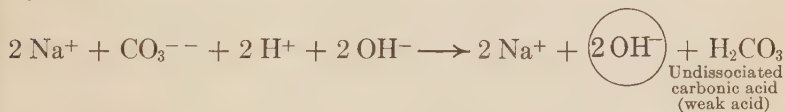
or even by synthesis, from its elements:



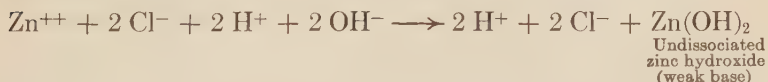
106. Names of salts. Salts are named after the acid from which they are derived:

ACID	NAME	SALT	NAME
H ₂ SO ₄	Sulphuric	Na ₂ SO ₄	Sodium <i>sulphate</i>
HCl	Hydrochloric	KCl	Potassium <i>chloride</i>
HNO ₃	Nitric	KNO ₃	Potassium <i>nitrate</i>
HC ₂ H ₃ O ₂	Acetic	Pb(C ₂ H ₃ O ₂) ₂	Lead <i>acetate</i>

107. Most salts are not neutral. Since they are formed by neutralization and since many of the common ones are neutral to litmus, it is sometimes thought that salts are neutral substances; but this is far from the truth. For instance, sodium carbonate in solution acts like a base. (It has a slippery feeling and turns litmus blue, etc.) Why? The explanation has been given under hydrolysis (par. 95), where it was shown that, according to the equation



the undissociated carbonic acid removes H⁺ from the solution and leaves it with an excess of OH⁻, giving rise, therefore, to a basic solution. Now this will be true whenever the acid formed is weak, that is, not highly dissociated. Also, if the base formed is weak, the solution will have an acid reaction; as, for instance:



whereas, when both acid and base are strong, both H⁺ and OH⁻ will be withdrawn from solution, and it will be neutral, thus, for instance, sodium chloride is neutral because it is derived from

sodium hydroxide, a strong base and hydrochloric acid, a strong acid. To sum up :

SALT WHOSE SOLUTION IS	FORMED FROM
Neutral	strong acid + strong base
Acid	strong acid + weak base
Basic	weak acid + strong base

In an approximate way, this enables us to predict the reaction of most salts.

COMMON STRONG ACIDS	COMMON STRONG BASES
HCl	NaOH
HNO ₃	KOH
H ₂ SO ₄	Ca(OH) ₂
	Mg(OH) ₂
	Ba(OH) ₂

We can easily remember these and assume others weak. That will mean, then, that

NEUTRAL SALTS	SALTS WHOSE SOLUTIONS ARE ACID	SALTS WHOSE SOLUTIONS ARE BASIC
Sulphates } { Na	Sulphates } { of	Na } { Salts
Chlorides } { K	Chlorides } { any	K } { of
Nitrates } { Ca	Nitrates } { other	Ca } { acid
		Mg
		Ba

Let us test this idea : What effect will mercuric chloride solution, HgCl₂, have on litmus? It is a salt of HCl, a strong acid ; but the base Hg(OH)₂ is not one of the strong bases ; so the solution will turn litmus red. Substances whose solutions are basic are called

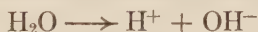
alkalies, and these include the strong bases (NaOH, etc.) as well as their salts, such as Na_2CO_3 , K_2CO_3 , $\text{Na}_2\text{B}_4\text{O}_7$, Na_2HPO_4 , etc.

108. Kinds of compounds found in nature. It seems reasonable to suppose that there cannot be many active substances found in nature, because as soon as they come into contact with something else a chemical change takes place. For that reason, we do not find large amounts of bases or acids. These may sometimes be present in plants, but not for any length of time. The rocks are almost entirely composed of salts, some soluble but neutral, others insoluble. Besides the silicates, we find NaCl , KCl , NaNO_3 , MgSO_4 , CaSO_4 , which are soluble and neutral, as well as large numbers of salts which are insoluble and which, therefore, do not hydrolyze, like PbS , PbSO_4 , $\text{Ca}_3(\text{PO}_4)_2$, AgCl , BaSO_4 , ZnCO_3 , ZnS , and CaCO_3 .

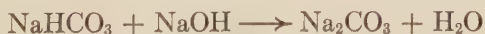
OPTIONAL PARAGRAPH

109. Neutralization in the human body. Every organ of the human body is markedly affected by changes in acidity or alkalinity. If the blood is too acid or alkaline, the heart will not pump properly. The acidity must be just right: a slight change in the acidity or alkalinity produces a vital change in the action of the heart. And this is true of all organs.

Now in the last chapter we found that a neutral solution contains equal concentrations of H^+ and OH^- ; which is true, of course, of water itself. Water, as you already know, slightly dissociates:



There are present in the blood certain substances, called buffer salts, which have a very useful function. If acid is added, they partially neutralize it, while if alkali is added, they can neutralize that, too. Among these substances are disodium hydrogen phosphate, Na_2HPO_4 , sodium bicarbonate, NaHCO_3 and proteins. For instance, if some strong alkali, like NaOH, is added to the blood, we may get



the strong alkali, NaOH, being reduced to a weaker alkali, Na_2CO_3 . On the other hand, if a strong acid, like HCl, is added, we may get



the strong acid HCl being converted to the neutral salt NaCl. In this way the body protects itself, by the use of a "defense mechanism," against the destructive effects of excessive acidity or alkalinity.

Table of Common Salts (Water of Hydration Omitted)

(Formulas need not be memorized)

COMMON NAME	CHEMICAL NAME	FORMULA	ONE USE
Alum	Potassium aluminum sulphate	$\text{KAl}(\text{SO}_4)_2$	Water purification
Baking soda	Sodium bicarbonate	NaHCO_3	Baking powder
Blue vitriol	Copper sulphate	CuSO_4	Batteries
Borax	Sodium tetraborate	$\text{Na}_2\text{B}_4\text{O}_7$	Cleaning
Chalk	Calcium carbonate	CaCO_3	Polishing powders
Chile saltpeter	Sodium nitrate	NaNO_3	Fertilizer
Copperas	Ferrous sulphate	FeSO_4	Making ink
Cream of tartar	Potassium acid tartrate	$\text{KHC}_4\text{H}_4\text{O}_6$	Baking powder
Epsom salt	Magnesium sulphate	MgSO_4	Laxative
Glaubers salt	Sodium sulphate	Na_2SO_4	Laxative
Hypo	Sodium thiosulphate	$\text{Na}_2\text{S}_2\text{O}_3$	Photography
Lunar caustic	Silver nitrate	AgNO_3	Caustic
Plaster of Paris	Calcium sulphate	CaSO_4	Casts
Rochelle salt	Sodium potassium tartrate	$\text{NaKC}_4\text{H}_4\text{O}_6$	Laxative
Sal ammoniac	Ammonium chloride	NH_4Cl	Batteries
Saltpeter	Potassium nitrate	KNO_3	Gunpowder
Sal soda	Sodium carbonate	Na_2CO_3	Washing powder
Saleratus	Potassium bicarbonate	KHCO_3	Baking powder
Table salt	Sodium chloride	NaCl	Manufacture of HCl
Washing soda	Sodium carbonate	Na_2CO_3	Same as sal soda
White vitriol	Zinc sulphate	ZnSO_4	Batteries

QUESTIONS

1. How could citric acid be made, using the general method for making acids?
2. Why are most acids sour? Why are they not all sour?
3. Some acids do not turn litmus red. Explain.
4. Write ionic equations for:
 - a. Milk of magnesia + hydrochloric acid.
 - b. Ammonium hydroxide + nitric acid.
 - c. Potassium hydroxide + sulphuric acid.
 - d. Calcium hydroxide + nitric acid.
 - e. Sodium hydroxide + sulphuric acid.
5. Define acid, base, salt.
6. Name two common bases, giving formulas and uses.
7. What is meant by an "alkaline reaction"?
8. Why does a solution of borax turn litmus blue?
9. What reaction would a solution of sodium bicarbonate have? ferrous sulphate? potassium nitrate? magnesium chloride? copper chloride?
10. What substance might be used as a paint remover?
11. Most of the compounds in nature are of what class? Explain why this is so.
12. Are all acids liquids? Explain.
13. In some old textbooks, CO_2 is called carbonic acid. Correct this statement.
14. If acid were spilled on a person's clothing, what should be used to prevent destruction of fabric? Why?
15. If strong alkali has been spattered over hands and face, what treatment would you apply? Why?
16. Why is some soap unfit for use on the face? What simple test could be made to ascertain this fact?
17. Why do we use boric acid rather than sulphuric acid as an eye wash?
18. In paragraph 108, name the salts mentioned.
19. Starting with the metal calcium, write equations for two methods of making calcium hydroxide.
20. Describe an electrolytic process for the preparation of potassium hydroxide.
21. Show by means of an ionic equation why a solution of ferric chloride gives an acid reaction.

OPTIONAL QUESTIONS

22. Why does a solution of borax, $\text{Na}_2\text{B}_4\text{O}_7$, turn red litmus blue? Explain this by means of an ionic equation.

23. Make electron diagrams of NaOH and HCl , and explain what takes place when they unite.

24. Suppose some hydrochloric acid were spilled on the street; Write an imaginary story telling what might happen to it until it reaches the ocean.

25. Name some substances other than litmus which change color with acids and bases.

26. Many people suffer from hyperacidity (excess acid) of the stomach. What substance would you suggest as a remedy? Write an equation to explain the use of this substance.

27. What is the probable source of the Cl^- ion in the HCl of the gastric juice?

28. What is a "buffer salt"? Explain how it works.

29. Outline a process for making soap.

CHAPTER XII

CARBON. CARBON DIOXIDE. CARBON MONOXIDE

CARBON

110. Carbon not only forms many compounds, but, what is more important, it is a constituent of animal and plant tissues, and a constituent of all the foods we consume. In the form of diamond and graphite, pure carbon is found in the free state. Coal contains large quantities of free carbon. The carbon in coal is *amorphous*, that is, it has no crystalline form; graphite and diamond, on the other hand, are crystalline varieties.

111. The diamond. Not only is this much-prized element found in nature, but as early as 1890 Moissan, the French chemist, succeeded in obtaining it artificially, though at such prohibitive cost, and the diamonds so small, that this method of manufacture has never become a commercial success. The principle of Moissan's method was to heat carbon in iron to an extremely high temperature in the electric furnace, and then plunge the mass into water. The pressure generated upon the interior portions by this sudden cooling converted some of the carbon into diamond. (Quite recently a number of scientists have cast some doubt upon this method.)

The diamond is not only a much-prized gem, but, being the hardest substance known, it is widely used for drilling purposes and as an abrasive.

112. Graphite is found in New York State, Ceylon, and other places; and is also made artificially by heating hard coal to a high temperature in the electric furnace. Mixed with clay, it is used

in the manufacture of lead pencils. It also finds use as a lubricant (being known as "oildag" if mixed with oil and "aquadag" if mixed with water); as a paint in order to prevent iron from rusting; and in the manufacture of crucibles because of its infusibility.



FIG. 100. Lead pencil made of graphite.

Graphite is sometimes called *plumbago* and sometimes *black lead*.

It should be noted here that lead pencils do not really contain lead at all, but graphite. The pencils are mixtures of graphite and clay, the hard pencils containing more clay than the soft ones.

113. Quite a number of substances contain varying amounts of amorphous carbon — the carbon with no crystalline structure. The following are some of them: coal, coke, wood charcoal, bone-black, and lampblack.

114. Coal is a form of fossilized wood — wood which has lain buried underground for centuries and centuries and lost much of its hydrogen and oxygen. In Chapter III you have seen how this planet of ours was formed and how in cooling the three portions, solid, liquid (water), and gas (the atmosphere), were separated out. You have also learned that the air, at present, contains a small per cent of carbon dioxide; but



FIG. 101. Pouring molten metal from a graphite crucible.

we have evidence to show that in the beginning there was much more, possibly as high as 40 per cent carbon dioxide. Now it is this carbon dioxide which the plants utilize and out of which they make their tissues, wood, and sugars, starches, and proteins. Small wonder, then, that plants flourished to such an extent that at one time the land was covered with a dense growth. The plants abstracted from the air an enormous amount of carbon dioxide and an equal amount of

oxygen was returned. Consider the equation usually given to represent this reaction :



The carbon dioxide decreased while the oxygen increased until animal life was possible ; and then began a balancing process in which the animal gave carbon dioxide to the plant, which in turn furnished oxygen to the animal. This is what is going on to-day.

Now much of this vegetation was submerged and ultimately covered with mud and sand and other *débris*. Probably due to the heat of the earth, chemical changes began to take place ; and since the earth was covered, the oxygen could not take part in these

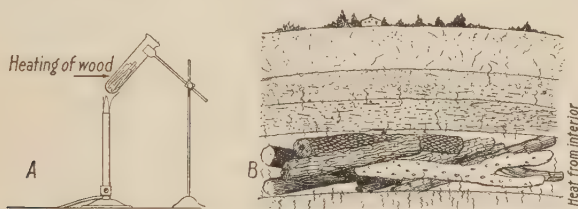


FIG. 102. The formation of coal is similar to heating wood in a test tube without air.

changes, which therefore must have been similar to what would happen if a piece of wood were heated in a test tube without air. That similar products were formed we have plenty of evidence.

The evidence for this theory of coal formation is overwhelming. There is often found the vein structure of leaves in coal. Trees have been found only partially carbonized, one end coal, the other still wood. With coal deposits we find marine animals as well as sedimentary rocks (formed in water). The heating of wood in a closed vessel forms gases such as hydrogen, methane (CH_4), and carbon monoxide (CO) ; liquid, like oils ; and a solid, charcoal. The natural gas usually associated with coal deposits contains methane and carbon monoxide, the hydrogen having probably escaped because of its great power of diffusion. Petroleum is, without doubt, the oily product, while the coal resembles charcoal in composition.

All these substances are found in such relation in the earth as to make it certain that they are derived from a common source. These changes have taken long periods of time, and are still going on.

115. Varieties of coal. We have at the present time many varieties of coal deposits in all possible stages of disintegration. Considering wood as $C_6H_{10}O_5$, there has been a gradual elimination of hydrogen, oxygen, and some carbon to form water, methane, and carbon monoxide, with a general tendency towards an increase in the percentage of carbon in the remaining product.

The following are some of the varieties of coal :

PRODUCTS	DESCRIPTION	APPROXIMATE		
		%C	%H	%O
Peat	Brown mass of leaves and roots	60	5	30
Lignite	Brown coal	65	5	20
Bituminous coal	Black, soft coal	75	5	10
Anthracite coal	Hard, dense, shiny coal	90	2	2

These products contain, in addition, small quantities of nitrogen sulphur, and so forth, and variable quantities of ash.

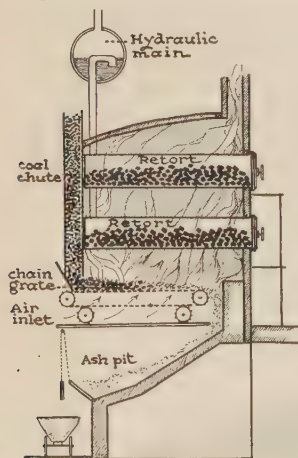


FIG. 103. Formation of coke.

Hard coal (*anthracite*) contains a large percentage of free carbon, and soft coal (*bituminous*) contains less free carbon and more carbon combined with hydrogen.

116. Coke. By heating coal in the absence of air (*destructive distillation*) a residue is left which is known as coke, containing much uncombined carbon, together with some mineral matter. It is used as a fuel, and, to an enormous extent, as a reducing agent in the preparation of iron from iron ore.

Just as coke is manufactured from coal by destructive distillation, so, by a similar process, *wood charcoal* can be obtained from wood and *boneblack* or animal charcoal from bones. The latter finds much use in the refining of sugar, and, in general, in removing coloring matter from certain substances. *Lamp-black*, made by burning oil or gas with an incomplete supply of air, is used in printer's ink.

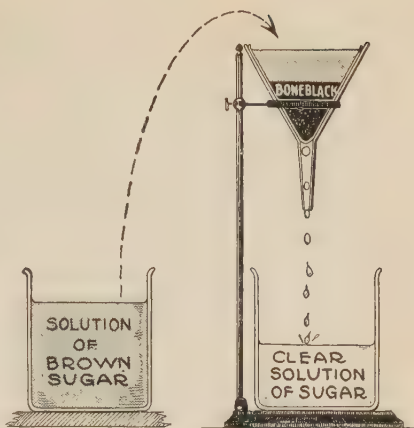


FIG. 104. Removing the color from sugar by the use of boneblack.

During the war, charcoal made from fruit pits was used in gas masks to adsorb poisonous gases.

117. General properties of carbon. Carbon is an extremely inert substance at ordinary temperature ; it is also highly insoluble. Charring a post preserves it for a very long time against decay. Carbon, like hydrogen, is a powerful reducing agent. It combines

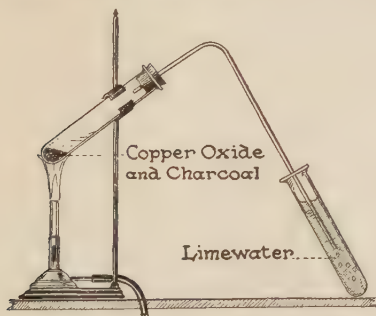


FIG. 105. Carbon used as a reducing agent.

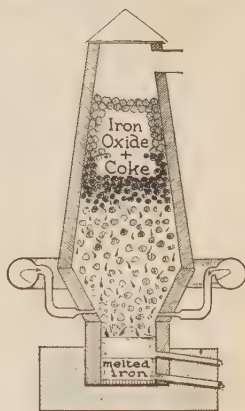


FIG. 106. Coke used as a reducing agent. (Scale 1" to 15').

with oxygen to form carbon dioxide. Copper oxide, for example, is reduced to copper not only by hydrogen (par. 45) but by carbon :



In a similar way, carbon, in the shape of coke, reduces iron oxide to iron. With metals, carbon forms carbides; one of these, calcium carbide, CaC_2 , is the source of acetylene; and another, silicon carbide or carborundum, SiC , is an excellent abrasive.

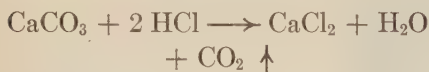
118. Allotropic forms. The diamond, graphite, and amorphous carbon are really different varieties of one and the same element, carbon; and this can be proved by the fact that when any one of these varieties is burned, we get carbon dioxide, and nothing else. In addition, just as we can convert amorphous carbon into diamond, so it is possible to convert amorphous carbon and diamond into graphite. The different varieties of one element, such as diamond, graphite, and amorphous carbon, are spoken of as *allotropic forms*. These differ in physical properties but are the same in chemical properties.

VARIETY	A CHARACTERISTIC PHYSICAL PROPERTY	ONE USE
Diamond	Very hard	Abrasive
Graphite	Very soft	Lead pencils
Hard coal	Shiny appearance	Fuel
Coke	Spongy appearance	Reducing agent
Lampblack	Powdery	Printer's ink
Boneblack	Powdery	Sugar refining
Wood charcoal	Light weight	Fuel

CARBON DIOXIDE

119. Whenever carbon or any compound containing carbon is burned, one product always formed is carbon dioxide, CO_2 . The breath which we exhale and one of the gases given off when coal is burned must, therefore, contain it. (You have probably blown through a tube into a test tube containing limewater, which then turns milky. This milkiness is due to carbon dioxide in our exhaled breath.)

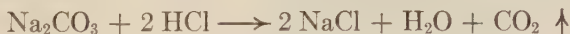
120. Preparation. Carbon dioxide is produced as a by-product in alcoholic fermentation, in limestone burning, and in the decay of animal waste; but for laboratory purposes a method originally devised in 1757 is used: the action of hydrochloric acid on calcium carbonate:



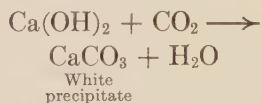
Much of the carbon dioxide used in soda fountains is made from the flue gases given off when coal is burned.

The flue gases are passed through a solution in which the carbon dioxide is soluble, and the solution is then boiled to liberate the gas. For the preparation of washing and baking sodas, where carbon dioxide is needed, the source is usually the burning of limestone.

121. Test for a carbonate. Any acid can be substituted for hydrochloric acid and any carbonate for calcium carbonate; for example:



In every case carbon dioxide is evolved, which can be tested with limewater:



Any carbonate may be substituted for the two given, with the same result. *If, then, an acid is added to a carbonate, carbon dioxide is*



FIG. 107. Preparation of carbon dioxide.

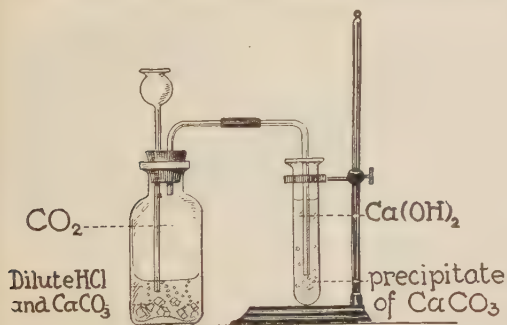
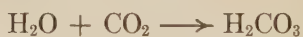


FIG. 108. Test for a carbonate.

evolved, and this gas can be tested with limewater.

122. Properties and uses. Carbon dioxide is a colorless gas, considerably heavier than air, and quite non-poisonous. We

would die in an atmosphere of carbon dioxide, not because the gas is poisonous but because of the absence of oxygen. Carbon dioxide is present in the air and is used by plants. It is fairly soluble in water, forming carbonic acid, a weak acid:



But this, in turn, is not a particularly stable acid, having a tendency to change back to CO_2 and H_2O . Some of it, however, dissociates into H^+ and HCO_3^- , thereby turning blue litmus red. So far no one has been able to isolate the acid H_2CO_3 .

All forms of aerated beverages, from sparkling Burgundy to Vichy, contain carbon dioxide.



FIG. 109. A fire extinguisher in action.

Carbon dioxide can be liquefied, and even solidified, without much difficulty. At a pressure of about 100 pounds per square inch the carbon dioxide gas can be converted into liquid. By rapidly evaporating the liquid, and by taking advantage of the heat absorbed thereby, the liquid may be converted into a "snow," which may then be compressed into dense blocks called "dry ice." By the use of this solid carbon dioxide, ice cream manufactured in New York and Philadelphia has been successfully transported to Cuba, and carloads of frozen fish have been shipped by rail on a five days' journey without re-icing.

Its heaviness and the fact that it does not support combustion make carbon dioxide useful as a fire extinguisher. One simple type of fire extinguisher is shown in Figure 109. When the ves-

sel is inverted, the solution of sodium bicarbonate and sulphuric acid mix, generating carbon dioxide, and the pressure of the gas forces water and gas out through the hose. In oil fires "foamite" is now frequently used; this consists of a mixture of aluminum sulphate and sodium bicarbonate:



The insoluble aluminum hydroxide, together with licorice extract which is added, forms a stiff "foam," which covers the surface of the oil.

In the raising (leavening) of bread carbon dioxide finds a very extensive use. The gas may be generated in the bread in one of two ways; either by the use of yeast or by the use of baking powder. When yeast is used, the carbon dioxide is released by the action of the enzyme in the yeast on the starch of the bread.

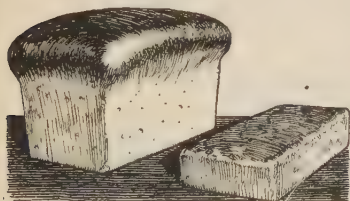
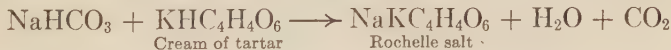


FIG. 111. Bread with and without leavening agent.

acid phosphate. If cream of tartar is used, then the reaction is:



Sometimes in the place of a tartrate, a phosphate salt is used.

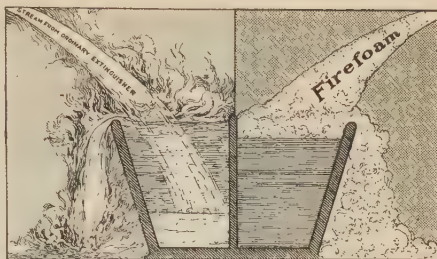


FIG. 110. An oil fire is not put out by water but is by Firefoam.

Carbon dioxide is also used in the manufacture of baking and washing sodas, and in the making of white lead.

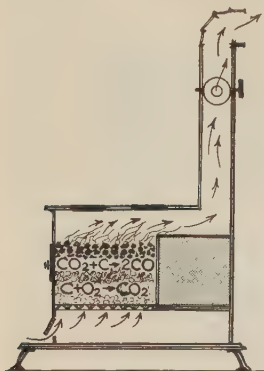
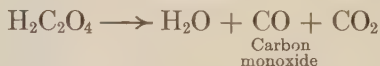


FIG. 112. Formation of carbon monoxide in a coal stove.

CARBON MONOXIDE

123. This gas is colorless, like carbon dioxide, but, unlike the latter, is inflammable and highly poisonous. The incomplete combustion of coal in the stove, often noticed by the burning of a pale, blue flame, is due to the formation of carbon monoxide; but in the laboratory the gas is most readily obtained by heating oxalic acid, $\text{H}_2\text{C}_2\text{O}_4$, with concentrated sulphuric acid:



The carbon dioxide is removed by passing the gases through lime water or sodium hydroxide. The sulphuric acid removes the elements oxygen and hydrogen from the oxalic acid in the form of water. This ability to extract water makes sulphuric acid a *dehydrating agent*.

124. Properties and uses.

Carbon monoxide, like hydrogen and like carbon itself, is a powerful reducing agent:

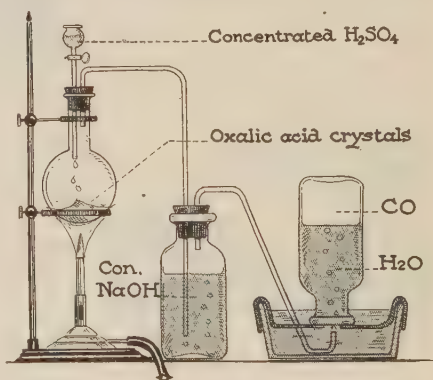
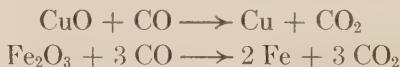
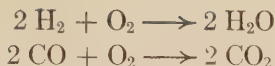


FIG. 113. Preparation of carbon monoxide.



and burns readily with a pale, blue flame similar to hydrogen, though the products in each case are different:



The carbon dioxide once formed can, of course, be identified by passing it into limewater.

Carbon monoxide is highly poisonous, due to its readiness to combine with the hemoglobin (the red pigment of the blood) and thereby prevent the oxygen from combining with the pigment. In combination with chlorine, phosphene, COCl_2 , is produced; and this highly toxic substance was used in the war as a poison gas.

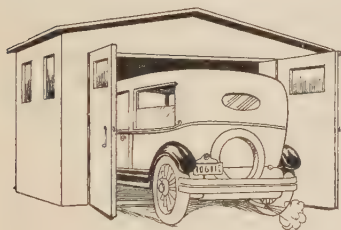


FIG. 115. The air in a garage may prove dangerous because of carbon monoxide.

D. H. Killefer, "Solid Carbon Dioxide," *Industrial and Engineering Chemistry*, vol. XIX (1927), p. 192.

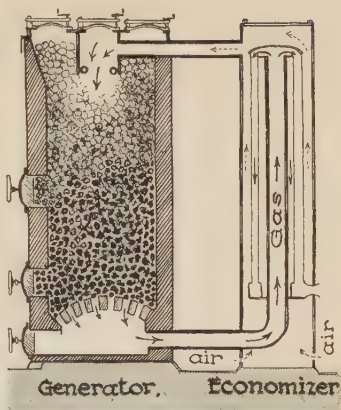


FIG. 114. Preparation of producer gas.

Carbon monoxide is a constituent of water gas and producer gas, two important fuels. The exhaust of automobiles contains this carbon monoxide—a source of constant danger.

SUPPLEMENTARY READING

A. D. Little, "The Romance of Carbon," *Industrial and Engineering Chemistry*, vol. XVIII (1926), p. 444.

QUESTIONS

1. Arrange a table showing the allotropic forms of carbon with one property characteristic of each variety and one use depending upon that property. (Do not copy the table in the text.)
2. What would you say is the outstanding chemical property of carbon? Illustrate your answer by at least two examples.
3. Describe a laboratory method for preparing carbon dioxide. Give the equation.

4. What purpose does the carbon dioxide serve (a) in a fire extinguisher, (b) in the baking of bread?

5. Describe a laboratory method for preparing carbon monoxide. Give the equation.

6. Write three equations to show that carbon, carbon monoxide, and hydrogen have one chemical property in common.

7. Given three bottles, one containing oxygen, the second carbon dioxide, and the third carbon monoxide, show how you would distinguish them.

8. Air containing relatively large quantities of carbon dioxide is not fatal; air containing relatively small quantities of carbon monoxide is fatal. Explain.

9. Given two bottles, one containing hydrogen and the other carbon monoxide, tell how you could identify each.

10. In the preparation of carbon dioxide and carbon monoxide, explain why the reactions go to completion.

11. Write equations to represent the complete combustion of marsh gas, CH_4 ; acetylene, C_2H_2 ; ethylene, C_2H_4 .

12. Foods contain carbon, hydrogen, and oxygen, among other elements. What products are formed when foods are burned in the body? Give evidence for your answer.

13. Name all the reducing agents you can and give equations to illustrate the reducing property of each

OPTIONAL QUESTIONS

14. Give a brief account of the principle used by Moissan to make the diamond artificially.

15. Write the ionic equation for the preparation of CO_2 , using sodium carbonate and sulphuric acid.

16. In the foamite process, as well as in the use of one form of baking powder, aluminum sulphate is used because of its acid properties. Explain, using an equation.

17. We have learned that reactions take place between ions. How do you explain, then, the reaction between hydrochloric acid and calcium carbonate, which is insoluble? •

18. Name four products formed when coal is completely burned.

19. There is no higher oxide of carbon than CO_2 . With this in mind, explain why carbon monoxide burns, whereas carbon dioxide does not.

20. Why are peat and lignite brown, whereas hard coal is black?

21. Assuming that the formula for gasoline is C_8H_{18} , write an equation to show the incomplete combustion as it takes place in an automobile engine.

CHAPTER XIII

THE HALOGENS

125. If we examine the periodic table (page 51), we shall find that fluorine, chlorine, bromine, and iodine are grouped together under one vertical column — an indication that these elements are closely related. This relationship will become evident as we proceed with the chapter. In the meantime, we may merely mention that they are monovalent, that they form compounds with hydrogen (HF, HCl, HBr, HI) and with many metals (for example NaF, NaCl, NaBr, and NaI). You will, of course, recognize one of these compounds, NaCl, sodium chloride, which is none other than common salt. The elements we are now discussing are called *halogens*, from a Greek word which means “salt-producing,” because one of their sodium compounds is actually our common table salt, and the other three sodium compounds (NaF, NaBr, NaI) resemble table salt in their properties.

126. Chlorine, the commonest and best known of these elements, was first obtained by Scheele as far back as 1774 by a method still commonly used in the laboratory, though he was of the opinion that the gas was a compound containing oxygen. Not till forty years later was it definitely shown by Davy that chlorine was really an element. For the source of this chlorine one would naturally look to sodium chloride, its commonest and most widely distributed salt; and we can actually get chlorine from salt. In fact, the electrolysis of brine (a solution of salt in water), though primarily used to prepare sodium hydroxide, also yields enormous

quantities of chlorine. But a more convenient laboratory source is hydrochloric acid, HCl. By the use of an apparatus such as the

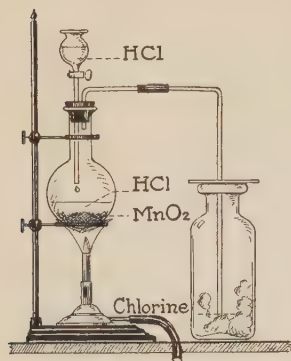


FIG. 116. Preparation of chlorine.

one employed in the electrolysis of water (par. 37) we can separate hydrochloric acid into its elements hydrogen and chlorine.

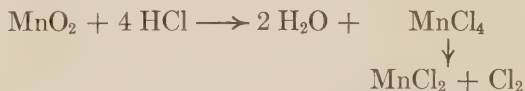
The usual laboratory method, however, depends upon the use of some oxidizing agent to remove the hydrogen in hydrochloric acid. Ordinary oxygen itself will do it, but rather slowly, though in the presence of a catalytic agent (in this case copper chloride) the reaction is quite vigorous :



In the laboratory, however, manganese dioxide is commonly employed as the oxidizing agent because it is cheap :



or, more correctly :



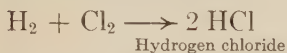
The manganic chloride is first formed, but being unstable, breaks up into manganous chloride and chlorine.

The manganese dioxide and hydrochloric acid are mixed in the flask as shown, the mixture heated, and the gas collected by downward displacement, since it is heavier than air and fairly soluble in water.

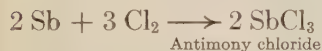
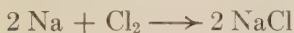
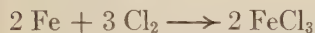
127. Properties. The gas has a very characteristic odor affecting the membranes of the nose and throat almost immediately. This odor, together with its light yellowish green color (its name is derived from a Greek word meaning "pale green"), makes it, as a

rule, easy to distinguish chlorine. It can be liquefied fairly easily, and as a liquid chlorine is used very extensively in the industries.

Though chlorine belongs to a different family of elements than oxygen, the two show certain similarities. For example, both combine with hydrogen :

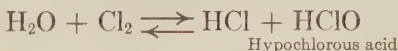


both combine with many of the metals :



These metals, when in a finely divided condition, burn quite brilliantly in chlorine ; so that the latter, like oxygen, may be looked upon as a supporter of combustion.

Chlorine first dissolves in water and then gradually combines with it :



If exposed to sunlight, the HClO decomposes :



Chlorine water, a powerful oxidizing agent, is a mixture of chlorine, hydrochloric acid, and hypochlorous acid and can therefore be kept without decomposing only in the dark.

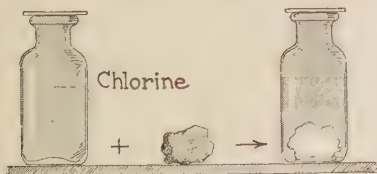


FIG. 118. Chlorine + colored cloth $\longrightarrow$ bleached cloth

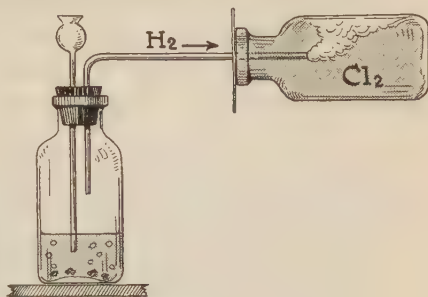


FIG. 117. Hydrogen burning in chlorine.

Chlorine is a powerful bleaching agent and disinfectant. Both of these properties are due not to the chlorine itself but to the hypochlorous acid. In bleaching it is the hypochlorous

acid which oxidizes the coloring matter of the cloth into a colorless compound ; dry chlorine shows no bleaching properties at all.

128. Uses. The bleaching property of chlorine is made use of in laundries to convert the yellow color of the linen into a white color. This is not objectionable unless large quantities are used; for then the fibers of the cloth are weakened. The chlorine used is in the form of bleaching powder, liquid chlorine, or Javelle water, which is NaClO (sodium hypochlorite).

Large quantities of chlorine, either as liquid chlorine or as bleaching powder, are used in the industries for bleaching purposes. The

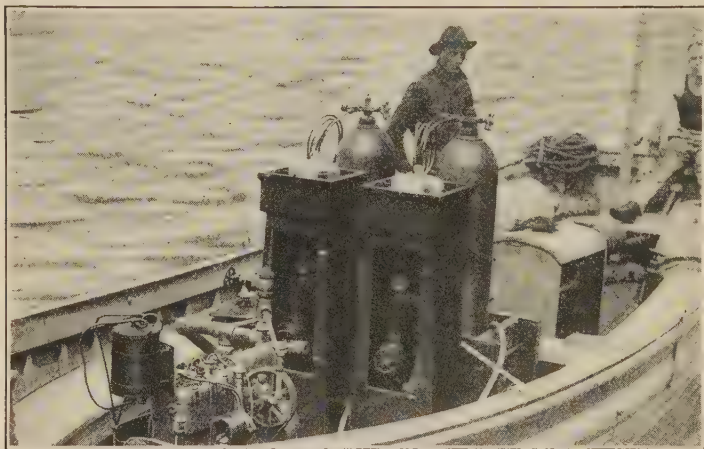


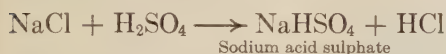
FIG. 119. Tanks of chlorine used to purify the water of a reservoir.

pulp and paper and the textile industries use vast quantities. As a disinfecting agent, chlorine is quite extensively used in water purification; and in the form of NaOCl (sodium hypochlorite) and more complex organic combinations is used very extensively for treating wounds and is known as *Dakin's solution*. Chlorine is also used in the preparation of chloroform and carbon tetrachloride and in the manufacture of a number of dyes and explosives.

Chlorine was the first poison gas used in the World War. Other poison gases such as phosgene (par. 124) and mustard gas contain chlorine.

129. Hydrogen chloride, as has already been stated, may be produced by the direct combination of hydrogen and chlorine — a method which is now being developed on a commercial scale. The hydrogen is obtained either by the electrolysis of water or from water gas; and the chlorine is a by-product produced in the electrolysis of brine. Incidentally, since in peace time more chlorine is produced than one finds a market for, this new method of making hydrochloric acid uses considerable quantities of chlorine which would otherwise go to waste.

In the laboratory, and on an industrial scale, we use common salt (NaCl) as the source of the chloride part of the molecule, and sulphuric acid as the source of the hydrogen part of the molecule:



The apparatus and method of collection are similar to that used for the preparation of chlorine.

At a higher temperature — one above that usually obtained in the laboratory — a second reaction may take place:



130. Properties. The gas, hydrogen chloride, is colorless, has a sharp odor, and is very soluble in water, forming hydrochloric acid (the “muriatic acid” of commerce). It forms dense white fumes with ammonia, a reaction sometimes made use of for identification purposes:



Hydrogen chloride is not an acid at all; it is the hydrochloric acid which is. The latter, like all strong acids, turns blue litmus red, is

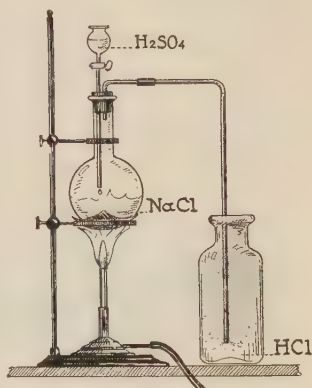
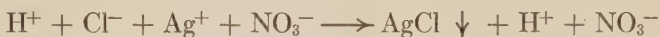


FIG. 120. Preparation of hydrochloric acid.

sour to the taste, and combines with bases to form salt and water; and like a number of other acids, is a good conductor of electricity — being highly dissociated — and liberates hydrogen when metals above hydrogen (see par. 44) are added to it. With silver nitrate it forms a white, curdy precipitate of silver chloride:



which is insoluble in nitric acid. This is used as a *test for any soluble chloride*. For example, a solution of sodium chloride will also react with silver nitrate to form silver chloride:

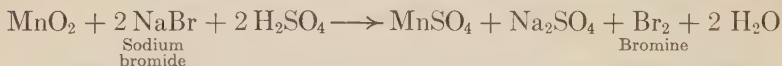


131. Uses. The acid is used to make ammonium chloride, now used so extensively in dry cells; to prepare glucose from corn starch; to extract glue from bones; to purify boneblack; cleaning metals before soldering or plating; etc.



FIG. 121. Hydrochloric acid used in soldering.

132. Bromine may be prepared in a way quite similar to chlorine: by the action of manganese dioxide on hydrogen bromide (hydrobromic acid). Since the latter, however, is not very stable, bromine is prepared in a slightly different way:



For purposes of convenience this reaction may be regarded as taking place in two steps: the first, the reaction of NaBr and H₂SO₄ to form hydrogen bromide, HBr; and the second, the oxidation of the HBr by the MnO₂.

133. Properties and uses. At ordinary temperatures, bromine is a very heavy, reddish brown liquid with a suffocating odor (the word "bromine" is derived from the Greek word meaning "stench"). It is slightly soluble in water, forming *bromine water*, which,

like chlorine water, is an oxidizing agent, though not so powerful in its action. As an oxidizing agent, bromine water finds extensive use in organic chemistry. A number of the bromides are used in medicine, and silver bromide is used in photography.

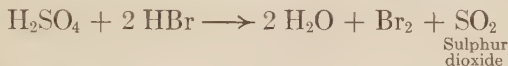
134. Hydrogen bromide (hydrobromic acid), like hydrogen chloride, may be prepared by burning hydrogen in bromine vapor:



or by reacting sodium bromide with sulphuric acid:

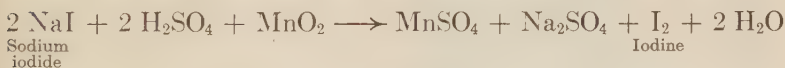


The instability of HBr gives rise to a secondary reaction:



so that besides HBr, sulphur dioxide and bromine may also be recognized among the products. Better results are obtained by using a somewhat diluted sulphuric acid, or by using phosphoric acid, H_3PO_4 in the place of sulphuric acid.

135. Iodine is prepared in the laboratory in an identical way:



though commercially the source is either seaweed (kelp) or sodium iodate, NaIO_3 , found in Chile saltpeter, NaNO_3 .

136. Properties and uses. At ordinary temperatures, iodine is a shining, purplish black solid. The vapor which is given off when iodine is prepared is violet in color (hence its name, which means "like a violet" in Greek); this vapor, when it condenses, goes over directly into solid iodine without the apparent formation, at first, of liquid iodine — a process which is known as *sublimation*. Like chlorine and bromine, iodine unites with hydrogen and with various metals; and in water solution shows oxidizing properties; but the reactions, in general, are feebler than those involving bromine, just as bromine, in turn, is less active than chlorine.



FIG. 122.
Use in medicine of salt containing bromine.

Iodine dissolved in alcohol is known as *tincture of iodine*, and is a common antiseptic. It is also very soluble in potassium iodide, forming a solution which finds extensive application in quantitative



FIG. 123. Iodine purified by sublimation.

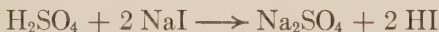
chemistry. Iodoform is another iodine compound which finds use as an antiseptic. A very interesting iodine compound is thyroxin, a complicated organic substance, which has been isolated from the thyroid gland of man and many animals, and which has even been synthesized in the chemist's laboratory. It would seem that the

function of this thyroxin in the body is to regulate the rate of oxidation; so that an excessive amount increases the oxidation within the cells to a very great degree, and the individual then becomes thin and nervous and jumpy. When there is less than the normal amount, the rate of oxidation decreases and the individual becomes fat and listless and heavy. Here much of the food, instead of being burned, is stored up as fat. Children having too little thyroxin may often be cured by adding a small quantity of sodium iodide to their diet.

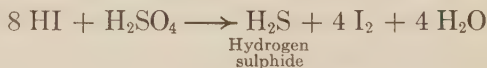


FIG. 124. Tincture of iodine.

137. Hydrogen iodide (hydriodic acid), like HCl and HBr, may be prepared by the action of sulphuric acid upon the corresponding salt:

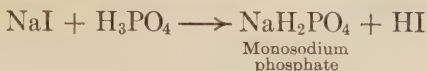


but the HI, being even more unstable than the HBr, gives little evidence of its presence, most of it reacting with more sulphuric acid:



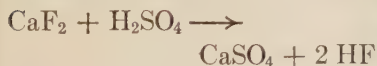
so that in the final products we get very little hydrogen iodide and considerable quantities of iodine and hydrogen sulphide.

Better results are obtained using phosphoric acid :



138. Fluorine, the most active of the halogens, remained unknown until Moissan finally isolated it in 1886. If hydrogen fluoride, HF, and manganese dioxide, MnO_2 , are heated, one would expect, reasoning by analogy, to get fluorine ; but such experiments always end in failure, for no sooner is the gas liberated than it attacks everything in sight, including the glass in which it is contained, to form compounds of fluorine. Moissan finally succeeded in getting the gas by electrolyzing anhydrous (without water) hydrogen fluoride, which, being a non-conductor, is first dissolved in potassium hydrogen fluoride, KHF_2 . Interestingly enough, it was shown that this fluorine, the most active of all the elements, does not even attack glass, provided both are perfectly dry — a reaction which vividly illustrates the important rôle moisture plays in chemical reactions.

139. Hydrogen fluoride, or **hydrofluoric acid**, is prepared by the action of calcium fluoride (fluor spar) on sulphuric acid :



This acid is of practical importance because it is used in etching glass, a reaction due to the fact that HF attacks the silica in glass :



the silicon tetrafluoride (SiF_4) so formed, being volatile and going off as a gas. In practice, the surface is coated with paraffin wax,



Keystone View Co.

FIG. 125. Etching of glass.

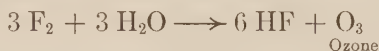
and the portions to be etched are exposed by removing the wax with a sharp instrument.

140. The halogens as a family. One property may be summarized thus:

F_2	Cl_2	Br_2	I_2
Gas	Gas	Liquid	Solid
Yellow	Greenish yellow	Dark red	Purplish black

With the exception of fluorine, they can be prepared by the action of manganese dioxide on the corresponding acid, or manganese dioxide and sulphuric acid on the corresponding salt. They show decreasing chemical activity as we proceed from left to right; which means that fluorine is the most active and iodine the least active. This can be illustrated in three ways:

(a) *The action of water.* Fluorine and water combine explosively:



Chlorine first dissolves in water and then gradually combines with it:



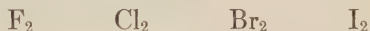
(Sunlight slowly decomposes $2 HClO$ into $2 HCl + O_2$.)

Bromine dissolves to some extent in water and remains as such indefinitely, unless exposed to the sun, when

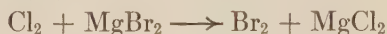


Iodine is but slightly soluble in water and is practically unaffected by it.

(b) *Displacement.* If we write the halogens again in the order



we may state that any element on the left will liberate the elements on the right of it when these are in a combined state; for example:



(But $I_2 + HBr$, no reaction.)

(c) *Action with hydrogen.* Fluorine combines rapidly with hydrogen even in the cold and in the dark; chlorine combines rapidly in the cold but only in strong sunlight; bromine combines only if heated; iodine combines but little even if heated.

The displacement method may be used to identify, for example, a bromide or an iodide; if a bromide, then the addition of chlorine water will liberate bromine, which, being very soluble in carbon disulphide or in chloroform, may be dissolved in one of these substances, yielding a very pronounced brownish red solution. If, on the other hand, we are testing for an iodide, then chlorine or bromine water may be added; in either case iodine is liberated, and this, on the addition of chloroform or carbon disulphide, yields a vivid violet solution; or the liberated iodine may be mixed with starch paste and a blue color will be obtained.

141. The halogen acids as a family.

HF

HCl

HBr

HI

are all colorless gases and may be prepared by the action of sulphuric acid on the corresponding salt; though, as has been pointed out, this method has its drawbacks when applied to HBr and HI. They are typical acids in the sense that they react with bases to form salts and with metals above hydrogen (par. 44) to liberate hydrogen. It is worthy of note that these acids (with the exception of HF) all react with oxidizing agents such as MnO_2 , H_2O_2 (hydrogen peroxide), HNO_3 , the halogen acids in turn being oxidized to water, the halogen being liberated. This is another way of saying that the halogen acids are reducing agents; and the reducing properties become more pronounced as we proceed from left to right.

QUESTIONS

1. Write equations to represent the preparation of chlorine (a) by the use of an oxidizing agent, (b) by a process of electrolysis, (c) by a method which, in principle, can also be used for the preparation of bromine and iodine.

2. Compare oxygen and chlorine as to their general chemical behavior.

3. If you burn a substance composed of carbon and hydrogen in chlorine gas, a black residue is left. What are the products?

4. Write equations to represent the combination of chlorine with (a) iron, (b) antimony, (c) potassium, (d) magnesium.

5. Write an equation to represent the bleaching action of chlorine.

6. What general method may be used for the preparation of the halogen acids? What shortcomings has this method?

7. Write equations to represent the reactions between (a) sulphuric acid and hydrogen bromide, (b) sulphuric acid and hydrogen iodide, (c) oxygen and hydrogen iodide.

8. To what use is hydrogen fluoride put? Represent this use by an equation.

9. Write a list of the halogens in the order of increasing activity. Write a list of halogen acids in the order of increasing stability. What very striking fact is brought out in these lists?

10. How would you distinguish (a) the halogens from one another; (b) the halogen acids from one another?

11. How would you test for (a) a bromide, (b) an iodide? Represent these tests by equations.

12. What is the iodine compound found in the body and of what value is it?

13. Why do you suppose liquid chlorine rather than chlorine gas is used for the purification of water and for bleaching?

14. Write ionic equations to represent the reactions between (a) silver nitrate and hydrogen bromide and (b) silver nitrate and hydrogen iodide.

15. How would you distinguish between (a) chlorine and oxygen; (b) bromine and hydrogen bromide; (c) iodine and graphite?

16. What other substance might be used in the place of manganese dioxide to prepare a halogen. Could this substance be used in the preparation of oxygen (potassium chlorate method)?

OPTIONAL QUESTIONS

17. In paragraph 140 (b), explain why any one of the replacement reactions can be written in the form $\text{Cl}_2 + 2 \text{Br}^- \longrightarrow \text{Br}_2 + 2 \text{Cl}^-$.

18. In the preparation of chlorine, why is it advisable to heat the mixture of manganese dioxide and hydrochloric acid by means of a water bath?

19. Why can chlorine be collected over a concentrated solution of salt (NaCl)?

20. Chlorine does not combine with metals if every trace of moisture has been removed. Suggest the possible function of water in such a reaction.

21. Write equations to show the preparation of bromine in two stages. (See par. 132.)

22. Write an equation to represent the probable reaction of phosphoric acid (H_3PO_4) and sodium bromide.

23. (a) Why is anhydrous hydrofluoric acid a non-conductor of electricity? (b) Why cannot an aqueous solution of hydrofluoric acid be used for the preparation of fluorine?

24. On exposure, hydrogen iodide turns brown. Why? Represent the reaction by an equation.

25. The order of activity of the elements we have studied may be given thus:

F Cl Br O I

Suggest an experiment to prove that oxygen belongs in the order shown.

26. What might you expect to get (a) when nitrogen and carbon dioxide are heated; (b) when carbon and carbon dioxide are heated? Explain. (Review question.)

CHAPTER XIV

THE COLLOIDAL CONDITION OF MATTER

142. The chemistry you have studied thus far has concerned itself chiefly with material of a scientific or technical interest. Very little has been mentioned of the substances we meet in our daily lives; our foods: bread, milk, butter, eggs, meat, cereals; our clothing: wool, cotton and silk fabrics, leather and rubber; our bodies with the materials and processes that make up life. These things have been neglected, because when it was attempted to

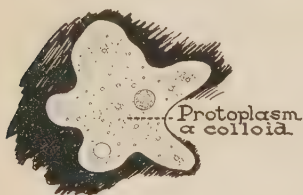


FIG. 126.

investigate them by the methods of "pure chemistry," they seemed very complex: they could not be obtained pure, they would not dissolve, and so little could be done with them. The reason for this, we now know, is that these substances do not readily break up into molecules, do not form true

solutions, and therefore, do not easily crystallize. They are in a condition called *colloidal*, in which the material has a new set of properties. Nearly every common thing we know and all living matter are in the colloidal condition. "Nature has chosen the colloid form in which to show her face."

143. What is a colloid? To define the colloidal condition in a few words is impossible. Professor Bancroft says: "Colloid chemistry may be called the chemistry of bubbles, drops, grains, filaments, and films." We get some idea from this definition: we are dealing with matter in a fine state of division, but not so fine as molecules in solution. Molecules in solution diffuse easily

through membranes like sausage skin, whereas colloids do not. There lies the difficulty in definition: we are dealing with material in an intermediate condition; in a transition stage between the finest state of subdivision, the molecular, and a coarse condition, like sand. In the following table the attempt is made to present this idea, together with some of the properties which accompany each state of fineness.

COARSE MATERIAL	MATERIAL IN COLLOID CONDITION	MOLECULAR
(25 mm. = 1 inch) Diameter $\frac{1}{10000}$ mm.	$\frac{1}{10000}$ to $\frac{1}{1000000}$ mm.	Smaller than $\frac{1}{1000000}$ mm.
	Degree of dispersion increases	
Held by filter paper	Passes through filters	Passes through filters
Easily seen by microscope	Not seen by microscope. Seen only by ultra-microscope	Cannot be seen by any microscope at present
No motion observed	Brownian movement. Does not dialyze	Molecular motion. Diffuses readily
Cloudy	Cloudy (usually)	Clear solution
Settles easily	Settles slowly	Does not settle
	<i>Colloids</i>	<i>Crystalloids</i>
Suspensions }	Suspensoids }	True solutions
Emulsions }	Emulsoids }	

It must not, however, be thought that there is a sharp line of distinction between these groups. As particles become smaller the time taken to settle becomes greater, cloudiness becomes gradually less, and so forth. The properties of matter change gradually with the size of the particle; and strictly speaking, it is not even the size of the particle, for sometimes one dimension may be large; as, for instance, in fibers and films. A fiber may be

long, but it is very thin. The important thing is the amount of surface compared to the volume, — the specific surface, as we say.



FIG. 127. Curdling of milk

Wool exposes a great surface per unit volume when compared with a rock: wool is in the colloidal condition. The rock may be made colloidal if it is sufficiently powdered. Milk is in the colloidal condition, consisting of particles of casein and fat floating in water.

It becomes less colloidal when this casein and fat is curdled into cheese, because the particles have collected together into larger masses presenting less surface. The same is true of blood when it clots, as is shown in the figure.

It should be understood, then, that colloids are not certain kinds of matter but rather a *condition* of material possible for any particular substance. It is not strictly correct to say salt is a crystalloid, protein is a colloid; for salt may be prepared in the colloidal condition, and some proteins may be crystallized; although it is true that we usually find salt as crystals and protein in the colloidal condition.

144. Properties of material in the colloidal condition. Colloids (that is, material in the colloidal condition) exhibit peculiar properties. These properties enable us to distinguish them rather easily. Each substance may not show all of these properties, but usually no room for doubt remains that the substance is colloidal. *In most of their properties, colloids lie between molecular and coarse material*, but some characteristics reach a maximum in the colloidal state, such as opalescence, color, and catalytic activity. Most gases and salt solutions are colorless, unless seen in enormous thickness, as in the atmosphere and ocean. Thin sheets of gold



FIG. 128. Clotting of blood a colloid.

appear green when held up to light. Colloidal gold may be red, blue, or green; colloidal silver: yellow, orange, red, violet, blue, green; all of these being due to particles of different sizes.

145. Turbidity. A colloidal solution usually appears cloudy, like milk or muddy water. The particles of suspended matter (casein and fat in milk, clay in water) are fairly large, and so scatter rays of light. Slight turbidity can be more easily seen by holding the solution against a black background and using sidewise illumination. Every one knows how easily dust particles are seen in a ray of sunlight entering a room with the shades drawn. Apparently visibility depends upon the difference in illumination between the



FIG. 129. Milk and muddy water are both turbid (characteristic of colloids).

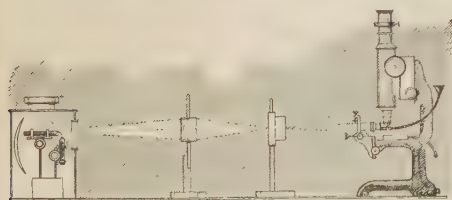


FIG. 130. The ultramicroscope.

object and our eyes. When the object is highly illuminated and our eyes are shaded, we see very well. The amount of light reflected from the object to our eyes is what counts. Now here is the entire principle of the *Tyndall effect* made use of in the ultramicroscope (Figure 130). Nearly all colloids show the Tyndall effect. If a true solution is subjected to a powerful beam of light from the side, it will remain dark. A colloidal solution shows the path of the beam of light because the particles reflect light, being nearly the same dimensions as wave lengths of light, whereas molecules, being much smaller, have no such effect.

The ultramicroscope is an ordinary microscope using the Tyndall principle to enable the observer to see still smaller particles. Some molecules have already been seen with the ultramicroscope and

there is no reason why with increased magnification and illumination we may not ultimately see all molecules.

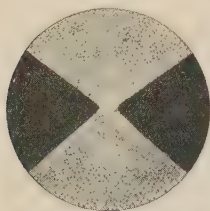


FIG. 131. Particles seen with the ultramicroscope.

This motion increases as the particles become smaller, as we should expect; until when the particles are of molecular size, their speed is identical with that called molecular motion.

147. Colloids pass through filters.

The separation of a solid from a liquid by filtration depends primarily upon the fact that the filter has small holes, smaller than the particles of solid, which are therefore held by the filter. In other words, a filter is a sieve. Ordinary filter paper has holes about $\frac{3}{1000}$ mm. (.003 mm.); the best filter paper, .001 mm.; unglazed porcelain filters, from .0002 mm. to .001 mm. Only the finest filters, therefore, can hold

back colloids, and it is not surprising that such materials will pass through ordinary filter paper. This often becomes exasperating

146. Brownian movement. The particles of a colloid viewed with the ultramicroscope are seen to be in motion, to oscillate in a zigzag fashion, as shown in the figure. This is called *Brownian movement*, after Robert Brown, the botanist, who first noticed it. Perrin has shown that this movement is the result of the blows given to the particles of colloid by molecules of the solvent.

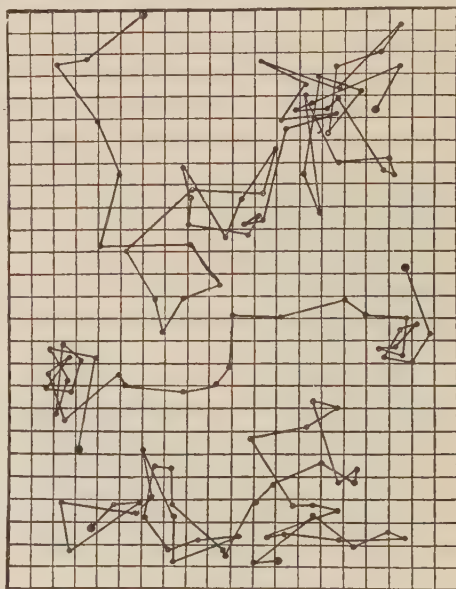


FIG. 132. Brownian movement.

during analytical work in chemistry; but it can be prevented by throwing down precipitates in more concentrated solutions or in the presence of an electrolyte, or sometimes by boiling. That the size of pore is not the entire reason for filtration may be shown by filtering a lamp-black suspension with and without a little soap. When the soap is added, the lampblack passes through the filter, although the particles have remained the same size, as can be shown under the microscope. Colloidal particles are electrically charged, and sometimes the paper fiber (which is also colloidal) is oppositely charged. These two unlike charges attract each other and so the lampblack is held. The soap, however, seems



FIG. 133. Filter paper fails to hold colloids.

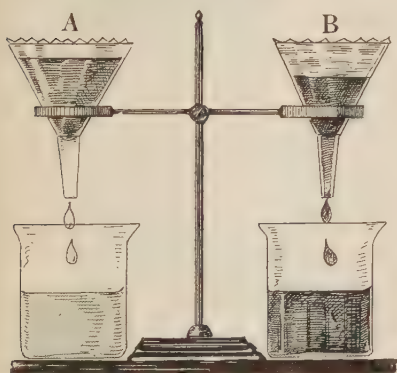


FIG. 134. Lampblack alone is held by filter; lampblack with soap is not.

to cover the lampblack and insulate its charge, allowing it to pass through. Colloids do not pass through most animal and plant membranes; nor do they pass through gels, like collodion; so that these substances may be used to filter them. Parchment, sausage casing, and collodionized filter paper are some of the materials used for "dialysis" or colloidal filtration. In this way colloidal solutions may be made more concentrated, for the dispersion

medium (solvent) passes rapidly through the filter, whereas the colloid is held. Hydrochloric acid on one side of parchment

paper reddens litmus on the other side in six seconds. Starch will not pass through such a membrane, but when it is changed to sugar by the enzyme ptyalin, it readily, dialyzes. This is what happens during the digestion of starches in the body.

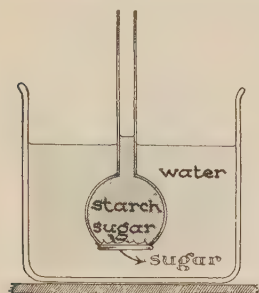


FIG. 135. Dialysis: a means of separating a colloid from a crystalloid.

148. Colloids do not settle; that is, they settle very slowly, and during the time we observe them, we notice no settling. This is also due, of course, to the size of the particle, as well as to the density and viscosity of the medium. This is taken advantage of in making paints where a viscous medium is always used; and the pigment is ground very fine, to prevent settling. It is, of course, the Brownian movement which keeps the colloid afloat, although gravity does have its effect. Perrin has found increased



FIG. 137. The materials in milk do not settle except when curdled.

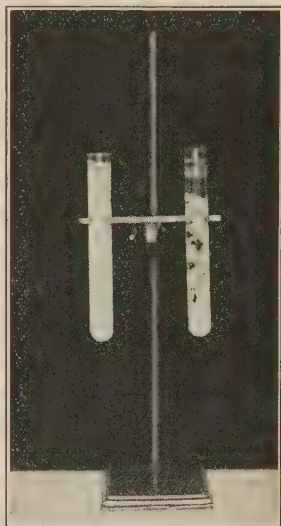


FIG. 136. White lead in oil is held in suspension, whereas white lead in water settles.

concentration of material at the bottom, in colloidal as well as in molecular solutions. It must indeed be a powerful force which will float platinum, in the colloidal condition, although the metal is twenty-one times as heavy as water.

149. Colloids are sticky and plastic. All sticky, tarry, plastic substances are in the colloidal condition. Glue, gum arabic, mucin-

lage, starch, flour, gelatin, asphalt, clay, soap, rubber, celluloid, cheese, and meat are a few examples. The adhesive property belongs to the colloidal condition.

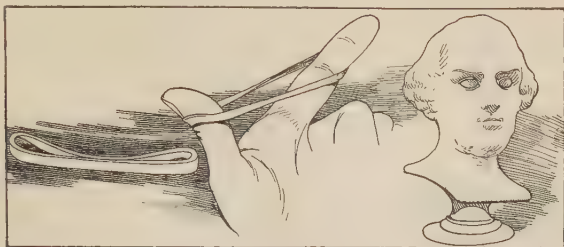


FIG. 138. Rubber is a plastic ; clay is a plastic.

The plastics, including such things as rubber, cements, clays, celluloid, and bakelite, are among our most useful materials in the trades ; and these will be taken up in a separate chapter. Plasticity corresponds to the gel stage, to be taken up presently, and is only an intermediate condition between the colloidal and the crystalline forms : the particles are constantly collecting into larger masses.

150. Foaming is a characteristic of colloids. It has been shown, by actual analysis, that, in a soap solution, there is more soap per cubic centimeter in the surface layers than in one cubic centimeter

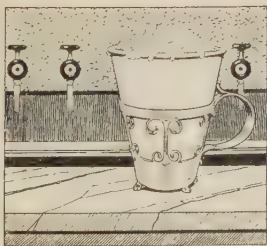


FIG. 139. Foaming.

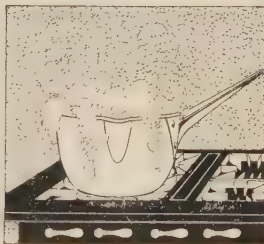


FIG. 140. Milk boiling over (foaming).

of any other part of the solution. "Foamite" is an application of a carbon dioxide foam for the purpose of extinguishing oil fires.

Foaming causes liquids to boil over easily, as may be seen when milk, cereals, and other foods are cooked. As has already been

stated, there is more material per cubic centimeter in surface layers than in the rest of the solution. This means that these layers contain less water and therefore dry out and burn more easily, while the rest of the solution is still liquid. For this reason, such foods are usually cooked in a double boiler.



FIG. 141. Cooking in a double boiler to prevent burning.

151. Swelling is a property of colloids with which we are all familiar. Beans, rice, and other foods, soaked in water, swell to many times their original size. Muscular and other tissues of the body frequently swell. Gelatin swells enormously in water. Rubber soaked in gasoline or carbon disulphide swells to many times its former size. What is the cause? The colloid is merely going into solution, and it is taking up the necessary liquid. The force with which it absorbs this liquid is enormous. The ancient Egyptians used to break rock by inserting a wooden wedge into a crack and then wetting the wood. Gelatin and water placed together in an iron cylinder will crack it. The process of water absorption is affected by the presence of small amounts of acid or alkali, a great swelling being caused by a slight increase in the concentration of H^+ (acid) or OH^- (base). The swelling caused by the sting of an insect is due to the presence of acid. The sting and the swelling may therefore be relieved by applying ammonium hydroxide to neutralize the acid.

When the housewife allows hard-caked cereal and milk pots to soak before cleaning them, she is redissolving the colloids by the absorption of water and the corresponding swelling.

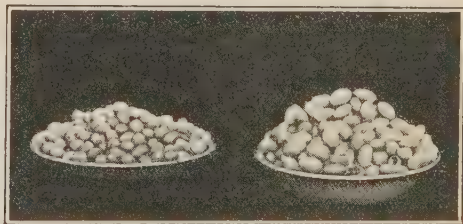


FIG. 142. Swelling of beans; before and after soaking.

152. Coagulation of colloids is a property familiar enough to all, but perhaps under different names. We have heard of coagulation of egg albumen, clotting of blood, setting of cement, curdling of milk, precipitation of silver chloride, cementing of metals, thickening of gelatin, congealing of meat juices, and stiffening of glue. Whenever we turn we find examples of *coagulation* or *gelation*, and in each case it has a different name. Gelation of a colloid means increased size of particle, loss of Brownian movement, and,

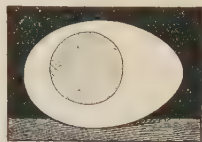


FIG. 143. A hard-boiled egg (coagulation).

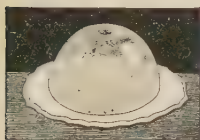


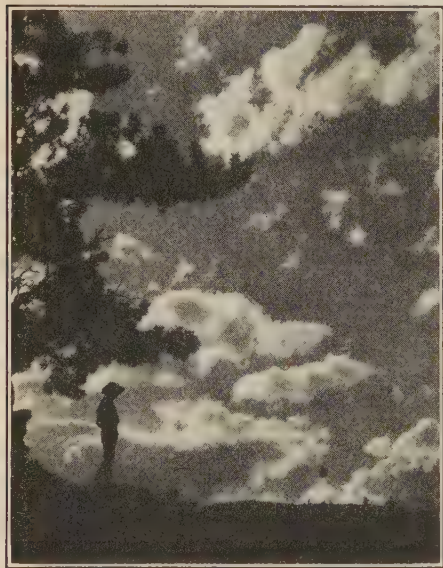
FIG. 144. Gelatin (coagulation).

ultimately, crystallization. *The opposite of gelation is called peptization.* Some colloids are referred to as reversible: gelatin coagulates at a low temperature and peptizes when warmed. Gelatin is a reversible colloid. Barium sulphate coagulates on addition of acid, but on washing out the acid, it goes back into solution.

153. Electrical charge on colloids. The electrical charge on colloidal particles has been mentioned before. Whenever matter is subdivided, like charges are left on the surfaces thus torn apart. Since the amount of an electrical charge depends upon the amount of surface, and since colloids show a great surface, they are highly charged. The charges on some common colloids are shown in the following table:

POSITIVE	NEGATIVE
Albumen	Starch
Hemoglobin	Mucilage
Hydroxides of Fe, Cu, Al	Silver chloride
Colloidal Pb, Fe, Cu	Silicic acid
	Oil emulsion
	Indigo
	Prussian blue
	Most powders like charcoal and kaolin

This charge causes the particles of the same solution to repel each other, thus giving rise, in part, to Brownian movement, and prevents their collecting together into masses and settling. A cloud, for instance, consists of charged particles of water. When a cloud of dust, oppositely charged, comes into contact with the water cloud, neutralization of the electrical charges takes place, and the



Erving Galloway

FIG. 145. Clouds: charged particles of water.

particles of water collect in drops and fall as rain. At the same time, all the little charges are frequently sufficient to make a lightning discharge.

Deserts probably owe their existence to the absence of clay in their soil. With nothing but rock and sand, what is there to overcome the electrical condition of the atmospheric water? Finely divided clay can assume a high charge, but coarse sand cannot. This may be shown very well by inject-

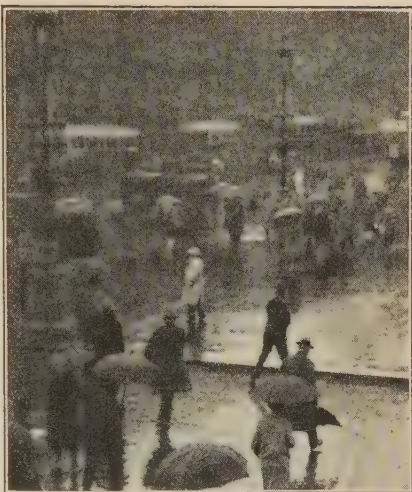
ing into an optically clear atmosphere (dust-free) a jet of live steam, when no cloud will be seen and no precipitation will occur. If into this atmosphere a little dust is now introduced, a cloud forms immediately and precipitation occurs at once.

The Cottrell System for preventing smoke and saving fine dust particles in gold ores depends upon the neutralization of their electrical charge. This method of precipitation has lately been applied to dispelling fogs with electrified fine sand, sprayed from airplanes.

To determine the character of the charge on a colloid, several

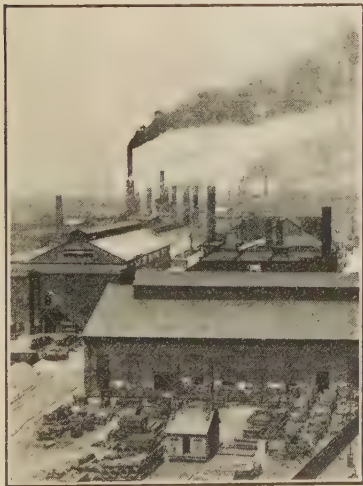
methods are in use, of which two will be mentioned. If we add to the unknown a solution containing a colloid of known charge, precipitation will be caused if the charge is opposite. Thus, knowing iron hydroxide, $\text{Fe}(\text{OH})_3$, to be positive and finding that it precipitates arsenic sulphide, As_2S_3 , we conclude the latter to be negative.

Another method — *electrophoresis* — is similar to electrolysis. When the current is turned on, all of the colloid migrates to one pole. Ferric hydroxide, for instance, collects about the negative



Keystone View Co.

FIG. 146. Rain caused by discharge of water particles.

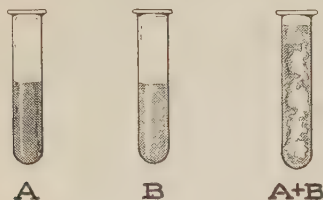


Ewing Galloway

FIG. 147. Two factories, the one on the right using the Cottrell system of smoke prevention.

pole. This may be used, in fact, as a method for concentrating colloidal solutions. For example, in purifying clay to make porcelain, the rust, $\text{Fe}(\text{OH})_3$, is separated from the clay, a negative colloid, by electrophoresis.

154. Classes of colloids. A colloidal solution resembles a true solution because there are at least two distinct substances



ferric hydroxide clear red
arsenic sulphide yellow
mixture precipitates

FIG. 148. Positive and negative colloids precipitate each other. See p. 141.

present. The solvent of the true solution is here called the *dispersion medium*, and the solute is called the *dispersed phase*. A colloidal solution is frequently called a *sol*. If water is the medium, it is called a hydrosol or simply a sol; with alcohol as a medium, the name becomes alcosol. There are

eight possible dispersed systems of two substances based on the physical states — solid, liquid, and gas :

	DISPERSED PHASE	DISPERSION MEDIUM	SPECIAL NAMES IF ANY	EXAMPLES
1.	solid	solid	solid solution: alloys	ruby glass; carbon in iron; brass
2.	solid	liquid	suspension	muddy water; paint
3.	solid	gas	smoke dust	smoke; air in flourmill
4.	liquid	solid		jelly; butter; zinc ointment
5.	liquid	liquid	emulsion	cream; milk; mayonnaise; cod liver oil
6.	liquid	gas	fog; cloud	steam in air; atmospheric fog
7.	gas	solid		pumice; meerschaum; gases in charcoal
8.	gas	liquid	foam	soda water; beer

The two most common dispersed systems are *suspensions* and *emulsions*. They differ markedly in one respect. A suspension may be easily coagulated, whereas an emulsion resists coagulation very strongly. Emulsions are thick or viscous. Examples are known of emulsions containing more than 99 per cent water which are so solid that they can be sliced with a knife. Emulsions are usually stable and reversible. They gel, giving off water; and when dry, swell and absorb water.

155. Precipitation and protection of colloids. Suspensoids may be precipitated by the addition of small quantities of electrolyte. The colloid always carries down with it the precipitating ion, for it is only one of the ions which is effective, namely, the one charged oppositely to the colloid. For instance, $\text{Fe}(\text{OH})_3$ is thrown down equally well by small quantities of Na_2SO_4 , MgSO_4 , and $\text{Al}_2(\text{SO}_4)_3$, whereas NaCl , MgCl_2 , and AlCl_3 have less effect. The precipitating effect increases with the valence. Emulsoids are, however, not very sensitive to this action, requiring quite concentrated solutions of electrolyte to cause gelation. Such substances are frequently used, therefore, to stabilize or protect a suspension; whence they are called *protective colloids*. Gelatin, glue, and gum arabic are powerful protective colloids. A little glue can prevent the setting of plaster of Paris for hours. It is the protective colloids in ice cream which

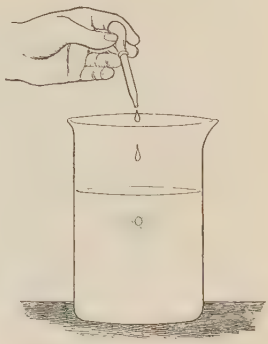


FIG. 149. Precipitation of a colloid by an electrolyte.

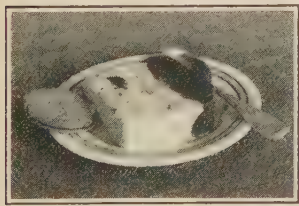


FIG. 150. Ice cream made "smooth" by gelatin (protective colloid).

prevent crystallization of the water into ice. In making candies, smooth texture depends upon the absence of crystallized sugar and this is brought about by the presence of gelatin and other

colloids. In analytical work, the presence of small amounts of protective colloids will slow down the chemical change so that a

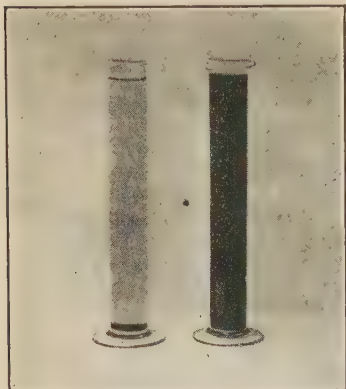


FIG. 151. Lampblack in water settles: lampblack in soap water does not settle.

precipitate may take hours or days to form, and so cause annoyance and error. Soap acts as a protective colloid in washing. This may be beautifully shown by shaking up some lampblack with water and again with soap water and allowing them to stand. The soap holds the lampblack in suspension, as shown by the uniformly black color of the solution. In washing with soap, the dirt particles are removed mechanically and then the soap, acting as a protective colloid,

holds them off the skin so that they are easily removed by water. The same principle is applied in using soap with steel wool for scouring aluminum pots. The soap removes the finely divided aluminum as soon as it is produced.

156. Other methods of gelation. Gelation may be brought about also by shaking and by centrifugalizing (rapid whirling). The centrifugal cream separator depends upon this action, while churning butter is an example of mechanical agitation. In general, acids and salts will cause coagulation, while alkalis cause peptization. Milk is curdled by acid, while ammonium hydroxide causes carbon to go into solution (as in India ink), and a little baking soda will soften the toughest dried peas or beans.

157. Syneresis. A gel gradually squeezes out a fluid or serum, which is a more dilute colloidal solution than the original. The material in going out of solution does not require so



FIG. 152. Churning butter to bring about gelation.

much solvent. This is the opposite phenomenon to swelling, in which solvent is absorbed. It is seen after the formation of blood clot (which yields blood serum), and after any colloid has been

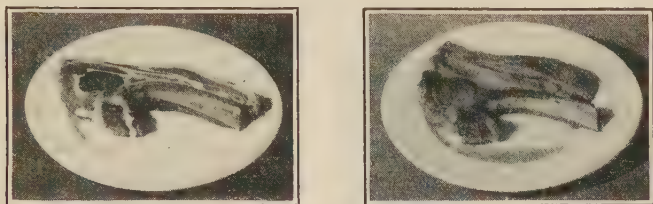


FIG. 153. A chop before and after broiling, illustrating syneresis.

coagulated for a time, as in the case of cranberry jelly. Whey is the serum of coagulated milk. The staling of bread is an example of syneresis. The moisture exuded may be detected when the bread is kept in a closed box: the water cannot evaporate and the outside of the bread becomes damp. Stale bread may be partly restored by warming the moistened bread in an oven, thus reversing the syneresis. When meat is heated, coagulation occurs, and this is followed by syneresis: the squeezing out of some of the juice.

158. Chemical activity of colloids. That colloids show greater chemical activity than the solid substance is not surprising, since interaction is proportional to surface of contact. Colloidal sulphur strongly reduces silver salts, whereas ordinary sulphur has no effect. A substance which seems inert may, in this way, be made quite active chemically; for, although it may be impossible to dissolve the substance, it can always be produced in colloidal condi-

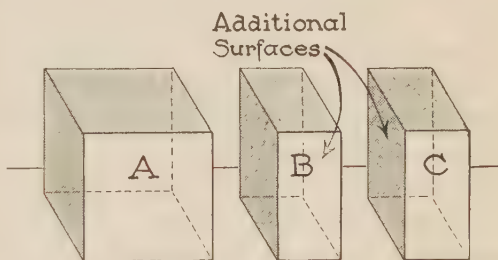


FIG. 154. Block A has been cut into B and C, producing additional surfaces.

tion. This surface action accounts for some catalytic action, as will be shown presently.

159. Adsorption. One of the most important results of the increased surface of colloids is *adsorption*. This is due to a con-

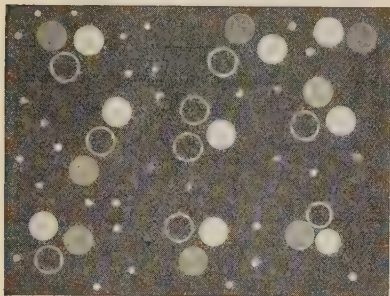


FIG. 155. Adsorption.

centration of material on the surface, shown in Figure 155. During the war, when a soldier wore a gas mask, he was depending upon the property of carbon to *adsorb* gases. This has been known for ages. Many sufferers from gas in the stomach eat charcoal tablets. Charcoal has also been used in ice boxes and in hospital wards

to remove gases (smells). Boneblack adsorbs many natural colors, which accounts for its use in decolorization of sugars and oils. Rubber adsorbs sulphur during vulcanization. Dirt particles adsorb soap and so are removed from a surface. Hygroscopic substances, like wool, adsorb water. Clay in soils adsorbs many soluble substances and so makes the soil fertile. Hemoglobin adsorbs oxygen. Practically all colloids show this property of adsorption. On account of the surface charge, a colloidal particle attracts and holds particles of an opposite charge. (See photo-micrograph, figure 155.)

In general, it is noticeable that the more complex substances are more readily adsorbed; we can therefore expect adsorption to be common among compounds found in the plant and animal kingdoms. Thus, the processes of dyeing, tanning, photography, treatment of rubber, cooking, and digestion show frequent examples of adsorption and

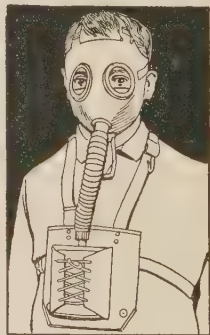


FIG. 156. Carbon used in gas mask to adsorb gases.

these will be taken up later. *Selective adsorption* is shown by kelp, which takes iodides out of sea water when the most painstaking chemical analysis reveals scarcely any. Carbon adsorbs ten times as much benzoic acid as acetic acid. Crystalline substances only adsorb others of an isomorphous character (*i.e.*, those which crystallize in the same shape as themselves). Calcium carbonate will adsorb sodium nitrate, both being hexagonal, but not potassium nitrate, which has a different shape.

Although heat will increase the rate of adsorption, it diminishes the total amount adsorbed. Some investigators think the process by which oxygen is carried to the tissues is, in part, adsorption.

The adhesive property of colloids is due to films adsorbed by the surfaces to be held together. If the film remains liquid, the mass is plastic, as in the case of quicksand.

The blue substance formed when iodine and starch come together is an adsorption compound. Even the familiar process of drying hands on a towel is adsorption. Adsorption frequently causes coagulation, since it concentrates the material. This same concentration on surfaces is frequently the cause of increased speed of chemical reaction, so that *all catalytic agents are adsorbents*. The adsorbent (catalyst) attracts to its surface both reacting substances which, brought into intimate contact, combine more rapidly than they would otherwise. Examples of this are shown in the manufacture of sulphuric acid by the Contact Process, with colloidal platinum as adsorbent; decarbonization of oils to make



FIG. 157. Drying one's hands: an illustration of adsorption.

gasoline, using colloidal nickel; and the manufacture of ammonia from nitrogen and hydrogen. *Catalysis*, then, is an increased rate of chemical change, brought about by a substance which itself does not undergo a permanent change during the process. A catalyst does not start a chemical reaction, but merely hastens one already going on, even if very slowly.

The enzymes, which influence so many physiological processes, are all catalysts. These are the result of cellular activity (life processes), and so far, they have not been produced artificially. Examples of such enzymes are ptyalin in the mouth, which changes starch to sugar; pepsin in the stomach, which acts on proteins; rennin in the stomach, which coagulates milk; and thrombin of the blood, which brings about clotting.

160. The colloidal and crystalline conditions. It is possible to prepare any substance either in colloidal or crystalline condition. Weimarn has succeeded in preparing many typical colloids, like albumen and aluminum hydroxide, in the form of crystals, while materials which are ordinarily crystalline, like common salt, can readily be made colloidal. It has long been known that most of the minerals exist in crystal form, whereas, when we produce such substances in the laboratory, they are gelatinous — colloidal. Diamond (carbon), quartz, heavy spar, sphalerite (zinc sulphide), galena (lead sulphide), and limestone occur crystalline in nature, but most of these substances are difficult to prepare in crystal form in the laboratory. When two reacting substances are brought together without regard to their concentrations, we get what we call a precipitate, which is at first a gel. Investigation has revealed the fact that if the solutions are dilute, we get small particles; so also with concentrated solutions. The largest particles are formed with a medium degree of concentration. Furthermore, the microscope has revealed the crystalline nature of these particles. In other words, *a colloid is composed of very tiny crystals*. The only difference between colloidal and crystalline material, then, is the *degree of dispersion, the size of particle, the development of surface*.

Time plays a very important rôle in the change from small to large crystals. That is why minerals are crystalline. Working on this idea, Weimarn has prepared crystals of barium sulphate and aluminum hydroxide (which are ordinarily colloidal) some of which took two years to form. Furthermore, there has been found to be a relation between solubility and crystal size. The more soluble substances form larger crystals in a given time.

We may now set down the stages through which a material passes from its formation to its final condition.

- a. Reacting ions brought together.
- b. Molecules form until saturation occurs.
- c. Molecules collect together until colloidal size is reached (opalescence).
- d. Gelation (microscopic crystals).
- e. Crystals form and continue to grow.

It will readily be seen that this change from *a* to *e* corresponds to a decrease in the degree of dispersion (or an increase in the size of particle), or a lessening of surface, resulting in a hardening of the particles.

161. Preparation of material in colloidal condition.

One essential condition for the stability of a colloid is that the material be insoluble in the dispersion medium. A colloidal solution of sodium chloride in water will not remain longer than the fraction of a minute. We may classify

methods of preparation as mechanical and chemical. In general, any process which will produce small particles in a medium in which they do not dissolve will make a colloid. In principle, this may be done in two ways: either by reducing a large mass to smaller masses or by increasing the size of particles in solution.

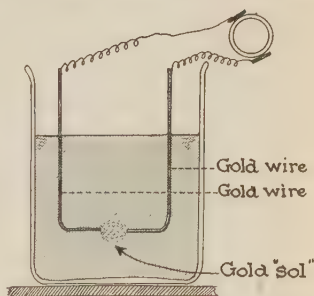


FIG. 158. Making a colloid by electrical dispersion.

An example of the first method is electrical dispersion of metals (Fig. 158); an example of the second method is the reaction between hydrogen sulphide solution and hydrogen peroxide to form colloidal sulphur.

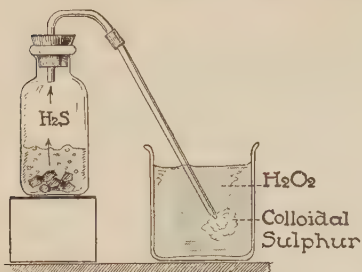


FIG. 159. Formation of colloidal sulphur by increasing the size of particles in solution.

162. Some practical applications of the principles of colloid chemistry. The principles of colloid chemistry explain many phenomena, and it is therefore essential that they should be understood by everybody. In the following paragraphs some

practical applications are given. Some of the topics referred to are discussed more fully in later chapters.

163. The colloidal condition in nature. The weather is frequently accompanied by changes in the colloidal condition. Fog and cloud are water in the colloidal condition; dust of the atmosphere is also colloidal. The interaction of these substances to precipitate rain and snow has already been referred to. The varied colors of the sky are due to the colloidal condition of the atmosphere.

The soil is a mixture of sand, clay, humus, and other substances in finely divided condition; in other words, a colloid. In fact, if earth has no material in the colloidal condition, it is sterile. Clay holds water which the plant may then take up slowly, and, in fact, it



FIG. 160. A delta caused by the precipitation of mud (a colloid) by salt (an electrolyte).



Sunset colors due to the colloidal condition of the air
(Facing page 150)

draws up water from the depths with a force comparable to the swelling of gelatin. It also adsorbs soluble material, thus preventing rain from washing out much of the fertilizing salts needed by plants. The power of streams to carry mud depends not only upon their velocity, but also upon the presence of protective colloids, like tannin, contained in the outer layer of most plants. When such a stream flows into a body of salt water, the coagulating effect of the electrolyte throws down a large part of this mud, which is therefore deposited as a delta. It should not surprise us, therefore, to learn that the largest delta, that of the Nile, has been formed by the particularly salty water of the Mediterranean.

Opal, chalcedony, and flint are quartz in the process of formation, — silica gels not yet crystallized. Agate is a silica gel, showing the rings characteristic of chemical action in a gel. The colors of many precious stones are due to colloidal hydroxides. Asbestos, talc, bog iron ore, and the clays are a few more examples of minerals in the colloidal condition.

164. Clothing. Cotton, wool, silk, linen, artificial silk, fur, leather, and rubber are material in the colloidal condition. They are particularly good adsorbents, holding fast water and many other materials like dyes. That is why they stain easily. Materials in the fibrous condition have remarkable tensile strength, some of them, like catgut, having a greater tensile strength per unit of weight than steel.

165. Photography. The coatings of plates, films, and photographic papers are composed of colloidal silver salts protected by gelatin. The latent image is an adsorption compound between

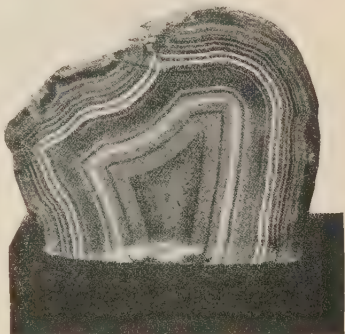


FIG. 161. Agate — the rings are due to precipitation in a colloid.

colloidal silver and silver salts of the halogens. For a fuller discussion of this subject see the chapter on Photography.

166. Cements. The whole process of the setting of cements is colloidal, and success depends upon an application of a knowledge of colloids. The ancient use of straw in making bricks implied the action of tannin to stabilize the colloid clay. White enamel is a



FIG. 162. Egyptians making brick.

suspension of stannic oxide, calcium phosphate, or some other white material in a glass.

167. Dyeing is chemical in its nature only when the same shade is produced with dilute as well as with concentrated solutions. But in most cases the more concentrated solutions produce deeper shades, and the process is therefore adsorption. Silk, in picric acid, is an example. The sulphur dyes all contain colloidal sulphur. The formation of skins on the surface of many dye solutions shows their colloidal nature. Mordants are colloids produced on fiber surfaces because they can adsorb dyes more strongly than the fiber. They therefore help to make the color "fast."

168. Tanning. This process is similar to dyeing. Tannic acid is used to coagulate the gelatin in the hide. With chrome tanning, a gel of chromium hydroxide is deposited. Before tanning, the hide is treated with acid so as to cause swelling; that is, to increase the degree of dispersion, so that the subsequent reaction with the tanning agent will be hastened.

169. Manufacture of soap. The materials from which soaps are made are nearly all colloidal. During the process the soap is

“salted out” or coagulated. This is not a common ion effect, as was at one time supposed, for it can just as well be accomplished by salts other than sodium salts.

All the processes making use of soap (washing, cleansing, etc.) involve colloidal changes.

170. Refining of sugar. The natural sugars contain gummy materials such as proteins in colloidal solution; and these are coagulated during refining by boiling with milk of lime. The natural coloring materials of sugars are removed (adsorbed) by filtration through boneblack.

171. Cellulose industries. All the applications of cellulose are colloidal in nature, since the raw materials, paper, cotton, and wood, are colloidal. Viscose is an emulsoid cellulose.

172. Cooking. Cooking involves the use of all kinds of colloids and therefore runs the entire gamut of colloidal changes. Souring milk shows syneresis. Starches swell in water; they are difficult to boil because they burn easily. Meat coagulates on heating. If soup is too salty, it may be made palatable by adsorbing the salt with raw egg. Coffee, tea, and cocoa are colloids.

173. Metallurgy. Alloys of which steel is a typical example are colloidal solutions of metals whose properties vary considerably with the degree of dispersion. According to Benedicks, the distinguished Swedish metallurgist, “The tendons of technology (steel cables) are, like the tendons of the human body, colloidal in structure.”

174. Colloidal fuels. These were developed in the United States during the recent war for use in atomizing burners. Mixed with oil and soap (protective colloid), these fuels offer great economy of space and produce very little ash. There is not so much danger with such fuels as with other oils which float on the surface; because here we deal with a fuel which is heavier than water. Should a fire break out, therefore, it can readily be extinguished with water. In fact the fire risk is as low as with anthracite coal.

175. Medicine. The tissues of the human body are colloids, and the various changes which take place in these tissues are colloidal in nature. It is possible, as Graham insists, that life itself is only a transition stage between colloid and crystalloid.

The enzymes which catalyze so many changes in the body processes are organic colloids. It has been suggested that œdema (swelling of tissues) may be due to increased adsorption of water in the presence of increased concentration of acid. Gold sols are being used in the examination of spinal fluid. Dr. Bell, of Liverpool, has suggested the use of colloidal lead in the treatment of cancer.

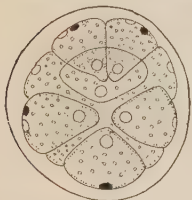


FIG. 163. The cell: a colloidal system.

The action of a muscle is caused by changes in its water content stimulated by lactic acid. The cell consists of protein, fats, and fat-like substances, all of which are colloidal. When it is called to mind that every living cell is surrounded by a membrane, and that membranes are colloidal in nature, the importance of colloid chemistry to physicians becomes apparent.

176. Miscellaneous. Ink is a colloidal solution of ferrous and ferric tannate, protected by gum arabic. Paints are suspensions of finely ground pigments in a viscous medium. Most paints are not quite colloidal, which causes their rapid settling out. Bakelite is a mixture of one colloid in its own gel. Rubber is colloidal and its changes are colloidal at nearly every step. Fresh rubber shows syneresis, squeezing out a serum rich in proteins. Vulcanization consists of an adsorption of sulphur and then, probably, the formation of an adsorption compound, ending with coagulation. Aquadag and oildag are colloidal graphites in water and oil.

Graham, the father of colloid chemistry, said: "The colloidal is, in fact, a dynamical state of matter, the crystalloidal being the statical condition. The colloid possesses energia. It may be looked upon as the probable primary source of the force appearing in the phenomena of vitality (life). To the gradual manner in

which colloidal changes take place may the characteristic protraction of chemico-organic changes also be referred." Graham believed, in other words, that since all living matter is in the colloidal condition, the duration of life corresponds with the time taken for the change from colloid to crystalloid. The crystallization of living material constitutes its death, and it has become certain that great discoveries regarding living organisms, both in health and disease, are to be found in an application of the principles of colloid chemistry to medicine.

SUPPLEMENTARY READING

- Emil Hatschek, *Introduction to the Physics and Chemistry of Colloids* (P. Blakiston's Son & Co.).
- Jerome Alexander, *Colloid Chemistry* (D. Van Nostrand Co.).
- W. M. Bayliss, *Principles of General Physiology* (Longmans, Green & Co.).
- C. W. W. Ostwald, *Introduction to Theoretical and Applied Colloid Chemistry* (John Wiley & Sons).
- W. D. Bancroft, *Applied Colloid Chemistry* (McGraw-Hill Book Co.).
- R. Zsigmondy, *Colloids and the Ultra-Microscope* (John Wiley & Sons).
- Hans Bechhold, *Colloids in Biology and Medicine* (D. Van Nostrand Co.).
- U. S. Bureau of Soils, Bulletin 52.

QUESTIONS

1. Name ten common substances which are usually in the colloidal condition.
2. State three properties characteristic of material in the colloidal condition.
3. Change the following into a more accurate statement: "Salt is a crystalloid."
4. What properties reach a maximum in the colloidal condition?
5. What is meant by dialysis? Give an example of its application.
6. In what two ways may settling be retarded?
7. What kind of liquids tend to boil over?
8. Why does a substance which sticks to the pot burn easily? How do we cook such foods?
9. Why are dried beans and peas soaked before cooking?
10. What is meant by a reversible colloid?
11. Mention examples of positive and negative colloids.

12. Explain how fine electrically charged sand may clear away fog.
13. Define "protective colloid." Mention an application of this principle, preferably one not mentioned in the chapter.
14. Explain the action of soap in washing.
15. State two methods of causing gelation.
16. A chop is served, and after standing awhile, a liquid appears on the plate. Explain.
17. What is adsorption? Mention a practical application of this principle.
18. Give common examples of emulsions and suspensions.
19. How can colloidal NaCl be prepared?
20. What means should be taken to soften hair in hides before removal? (Hair is a typical colloid.)
21. A piece of platinum has far less catalytic action than finely divided platinum. Explain.

OPTIONAL QUESTIONS

22. State three cases of coagulation not mentioned in this chapter.
23. Butter and cream are both colloids, containing water and fat. In what way do they differ?
24. Two liquids, like turpentine and alcohol, refuse to mix and separate into layers on standing. Suggest a means of assisting solution.
25. Mention applications of the principles of colloid chemistry in two industries, preferably not mentioned in the text.
26. Why must foods be digested before being absorbed?
27. Why will a small amount of the enzyme ptyalin, present in saliva, digest a very large amount of starch?

CHAPTER XV

CHEMICAL ARITHMETIC

177. Most of us purchase coal at so much per ton, and if we discover, after using it, that the quality has been inferior, we may change our dealer, but does that insure better coal? Does the dealer himself know which coal is good? And if so, how does he find out? This consideration is not limited to coal, of course.

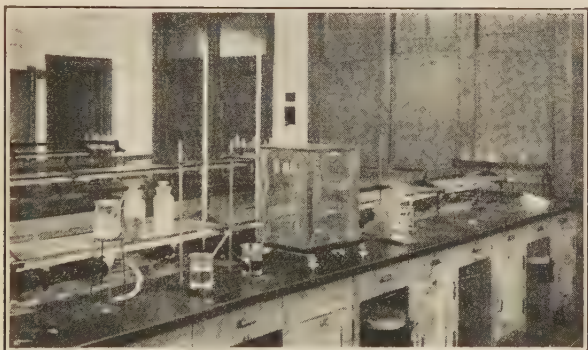


FIG. 164. A chemical laboratory where chemical analysis is done.

Suppose you are in the clothing business, and have ordered large quantities of woolens. When the cloth arrives, are you going to accept it on the basis of looks? When a large baking firm orders enormous quantities of flour, it must measure up to a sample, for the presence of a little more water will mean a loss of thousands of dollars, while other variations may affect the entire process of baking. These things are becoming more and more standardized, converting business from gambling into a scientific process. But how is it done? Chemical analysis is taking the place of guesswork.

Avogadro's hypothesis, then, may now be called a law. Working with his hypothesis, Avogadro was able to show — and this has since been confirmed by other methods — that the molecules of oxygen, hydrogen, nitrogen, and chlorine contain two atoms each. That means, if we call the atomic weight of hydrogen 1, the weight of the molecule must be 2. Now equal volumes of hydrogen and oxygen do not weigh the same, although they contain the same numbers of molecules. Why not? Because the molecule of oxygen weighs more than the molecule of hydrogen. Since one volume of oxygen weighs 16 times as much as one volume of hydrogen, one molecule of oxygen must weigh 16 times as much as one

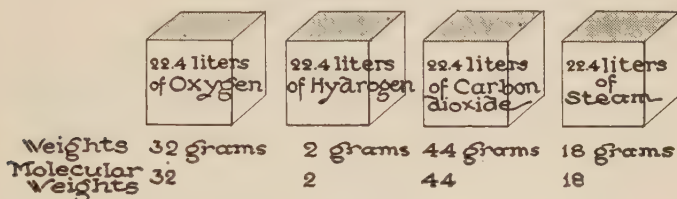


FIG. 165.

molecule of hydrogen, or 32 times as much as one atom of hydrogen. In other words the molecular weight of oxygen is 32. If we take 32 grams of oxygen gas, we find it occupies 22.4 liters. The same number of molecules of hydrogen gas is found in 22.4 liters of that gas, and this of course weighs 2 grams. In other words, weighing 22.4 liters of any gas gives us the *molecular weight* of that gas.

In the following problems you will see how useful is the idea of molecular weight in analytical chemistry. Where we deal with compounds other than gases, the molecular weight may be obtained from the formula by adding together an atomic weight for each atom in the molecule. For example, the molecular weight of Na_2CO_3 is 106 :

Na_2CO_3 which really means NaNaCOOO

and which therefore equals $23 + 23 + 12 + 16 + 16 + 16 = 106$

180. Percentage composition problems. Suppose we wish to determine the per cent of water in crystallized barium chloride. A sample is weighed, and then heated until all the water is driven out, when a final weighing is made. This tells us the weight of water in the sample, from which the per cent can be calculated. Let us assume the initial weight of the barium chloride crystals was 2.0000 grams, and the final weight, 1.7060 grams; the calculations, then, would be done as follows:

Weight of crystallized barium chloride	2.0000 g.
Final weight of barium chloride	<u>1.7060 g.</u>
Difference (<i>water</i> driven off)	0.2940 g.
Water, then, is, 0.2940 g. out of 2.0000 g.	

That is, $\frac{0.2940}{2.0000} = 0.1470 = 14.7\%$ *Ans.*

If, on the other hand, the formula is given, $\text{BaCl}_2 \cdot 2 \text{H}_2\text{O}$, and it is required to find the per cent of water, we proceed as follows:

Molecular weight of $\text{BaCl}_2 \cdot 2 \text{H}_2\text{O} = 244$

(atomic weights $137 + 71 + 4 + 32 = 244$)

Weight of water in the molecule = 36

The water, then, weighs 36 out of 244; $= \frac{36}{244} = 0.147 = 14.7\%$.

Ans.

The per cent of any of the constituents can be calculated in the same way:

Water, as above,	= 14.7%
Barium is 137 out of 244; $= \frac{137}{244} = .561$	= 56.1%
Chlorine is 71 out of 244; $= \frac{71}{244} = .291$	= 29.1%
check	<u>99.9%</u>

(Note that the total should check to 100%; but failure to carry figures to the proper number of decimal places may result in 99.9% or 100.1%.)

181. Problems on equations. It frequently becomes necessary to find out how much of a certain substance may be produced by a chemical change. These problems are of several types.

TYPE I. *Given*: a weight; *required*: a weight.

Example:

How much HCl could be obtained from 100 g. of pure NaCl ?

required wt. given wt.

- | Steps | Description | Actual Operations |
|-------|--|---|
| 1. | Write the equation. | $2 \text{ NaCl} + \text{H}_2\text{SO}_4 \longrightarrow 2 \text{ HCl} + \text{Na}_2\text{SO}_4$ |
| 2. | Mark given and required substances thus: $\square \square$, and forget the others. | $\boxed{2 \text{ NaCl}} + \text{H}_2\text{SO}_4 \longrightarrow \boxed{2 \text{ HCl}} + \text{Na}_2\text{SO}_4$ |
| 3. | Find the molecular weights of $\square \square$. | $ \begin{array}{l l} 2 \text{ Na} = 2 \times 23 = 46 & 2 \text{ H} = 2 \times 1 = 2 \\ 2 \text{ Cl} = 2 \times 35.5 = 71 & 2 \text{ Cl} = 2 \times 35.5 = 71 \\ 2 \times \text{mol. wt. NaCl} = 117 & 2 \times \text{mol. wt. HCl} = 73 \end{array} $ |
| 4. | Place mol. wts. above $\square \square$; and actual wts., given and required (x), below $\square \square$. | $ \begin{array}{c} 117 \qquad \qquad \qquad 73 \\ \boxed{2 \text{ NaCl}} + \text{H}_2\text{SO}_4 \longrightarrow \boxed{2 \text{ HCl}} + \text{Na}_2\text{SO}_4 \\ 100 \qquad \qquad \qquad x \end{array} $ |
| 5. | Make the proportion, reasoning as follows: if 117 g. NaCl were used, we would obtain 73 g. of HCl; therefore, if 100 g. NaCl are used, we shall obtain x g. HCl. | $ \begin{array}{l} \overbrace{117 : 73 :: 100 : x} \\ 117 x = 7300 \\ x = 62.4 \text{ g. HCl. } \textit{Ans.} \end{array} $ |

TYPE II. *Given*: a weight; *required*: a volume. Or *given*: a volume; *required*: a weight.

For this type of problem we shall have to make use of Avogadro's Law, and the gram molecular volume, 22.4 liters. Let us try this problem:

How many liters of chlorine could be obtained from

required volume

50 g. of pure NaCl ?

given weight

<i>Steps</i>	<i>Description</i>	<i>Actual Operations</i>
1.	Write the equation.	$2 \text{ NaCl} + 2 \text{ H}_2\text{SO}_4 + \text{MnO}_2 \longrightarrow \text{Cl}_2 + \text{Na}_2\text{SO}_4 + \text{MnSO}_4 + 2 \text{ H}_2\text{O}$
2.	Mark given and required substances thus: $\square \square$, and forget the others.	$\boxed{2 \text{ NaCl}} \longrightarrow \boxed{\text{Cl}_2}$
3.	Place given wt. below, and x for req. vol.; above given wt., mol. wt.; above req. vol., the gram mol. vol.	$\begin{array}{ccc} 117 & & 22.4 \\ \boxed{2 \text{ NaCl}} & \longrightarrow & \boxed{\text{Cl}_2} \\ 50 & & x \end{array}$
4.	Make proportion, reasoning as follows: 2 mol. wts. (117 g.) of NaCl will produce 1 mol. vol. (22.4 liters) of chlorine; therefore 50 g. of NaCl will produce x liters of chlorine.	$\begin{array}{l} 117 : 22.4 :: 50 : x \\ 117x = 11200 \\ x = 9.6 \text{ liters. } \textit{Ans.} \end{array}$

If the volume is given, and the weight required, the problem is done in the same way, x being the required weight in grams. Take note that if more than 1 molecular weight of the gas is concerned, then the volume will be 2, 3, etc. times 22.4 liters. Let us work a problem of this type.

Example: How much pure KClO_3 would be required to produce 100 liters of pure oxygen?

<i>Steps</i>	<i>Description</i>	<i>Actual Operations</i>
1.	Write equation.	$2 \text{ KClO}_3 \longrightarrow 3 \text{ O}_2 + 2 \text{ KCl}$
2.	Mark given and required.	$\boxed{2 \text{ KClO}_3} \longrightarrow \boxed{3 \text{ O}_2}$
3.	Place mol. wt. and gram mol. vol. above; real wt. and vol. below.	$\begin{array}{ccc} 245 & & 67.2 \text{ or } (3 \times 22.4) \\ \boxed{2 \text{ KClO}_3} & \longrightarrow & \boxed{3 \text{ O}_2} \\ x & & 100 \end{array}$
4.	Proportion.	$\begin{array}{l} 245 : 67.2 :: x : 100 \\ 67.2x = 24500 \\ x = \frac{24500}{67.2} = 365 \text{ g. } \textit{Ans.} \end{array}$

Example: 1000 c.c. of air at a pressure of 15 lbs. per square inch will become what volume at 60 lbs. per square inch?

P	V
15	1000
60	x

Proportion: $1000 : x :: 60 : 15$
[inversely]

$$15000 = 60x$$

$$x = \frac{15000}{60} = 250 \text{ c.c. } \textit{Ans.}$$

Charles discovered that a gas expands $\frac{1}{273}$ of its volume, at 0°C. , for each degree centigrade increase of temperature (pressure constant). He reasoned that if a gas lost $\frac{1}{273}$ of its volume for each degree below zero, then, at 273° below zero, it would have lost $\frac{273}{273}$ of its volume; that is, it would lose all its volume, and cease to exist. He therefore called this point -273°C. , the *absolute zero*. Now, Charles was not correct in this conclusion; because, as it cools, a gas changes to a liquid, which no longer obeys the law; but we may say, at any rate, that all gases must liquefy before they reach -273°C. , and that has been shown to be true. Charles constructed another thermometric scale, called the absolute scale, merely adding 273° to the centigrade readings, as follows:

CENTIGRADE	ABSOLUTE
20	293
10	283
0	273
-273	0

and we may now state Charles' Law: *The volume of a gas varies directly as the absolute temperature* (pressure constant). That

means, if the *absolute* temperature doubles, the volume doubles. A problem will make clear the use of Charles' law :

Example: 500 liters of hydrogen at 0° C. will change to what volume at 25° C.?

V	T (CENT.)	T (ABSOLUTE)
500 liters	0°	$(0 + 273) = 273$
x	25°	$(25 + 273) = 298$

Notice, we first find absolute temperatures by adding 273° to centigrade temperatures, and then make the proportion

$$x : 500 :: 298 : 273$$

$$\text{or } \frac{x}{500} = \frac{298}{273}$$

$$x = 500 \times \frac{298}{273}$$

$$x = 545.8 \text{ liters. } \textit{Ans.}$$

Problems in both laws together are worked as follows:

Example: A change in conditions from 20° C. and 750 mm. pressure to 0° C. and 760 mm. pressure would produce what volume out of 100 liters of oxygen?

V	P	TC	TA
100	750	20	293
x	760	0	273

We reason as follows:

The new volume, x , will be the present volume multiplied by two fractions, one for each law: $100 \times \text{---} \times \text{---}$.

First consider pressure: it goes from 750 to 760, becoming more; so that the volume must be less (inversely). One fraction, then, must be $\frac{750}{760}$, because this will make the volume less, whereas $\frac{760}{750}$ would make it more. So far then we have $x = 100 \times \frac{750}{760} \times \text{---}$.

For temperature, the change is from 293 to 273 (use only absolute readings). When temperature gets less, volume gets less (directly), so this fraction must be $\frac{273}{293}$, which will make the answer less; whereas multiplying by $\frac{293}{273}$ would make it more. We have, then, to sum up,

$$x = 100 \times \frac{75.0}{76.0} \times \frac{273}{293} \text{ liters. } \textit{Ans.}$$

183. To determine the formula of a gaseous compound. Suppose, for instance, we wish to determine the formula of marsh gas, one liter of which weighs 0.72 g. (at standard conditions), and whose composition is shown by analysis to be

$$\text{C} = 75\%, \text{H} = 25\%$$

The molecular weight of marsh gas is the weight of 22.4 liters, or $22.4 \times 0.72 = 16$. Now proceed as follows:

<i>Step</i>	<i>Description</i>	<i>Actual Operation</i>
1.	Find total weight of all the atoms of carbon in one molecule.	75% of 16 = 12
2.	Do the same for the other atom or atoms.	25% of 16 = 4

But since 12 is the total weight of all the atoms of carbon in this molecule, there can be only one atom of carbon (since the atomic weight of C = 12). Likewise, 4 is the total weight of all the hydrogen atoms in one molecule; and since each atom of hydrogen weighs 1, there must be 4 atoms of hydrogen. Therefore the formula of marsh gas is C_1H_4 or CH_4 .

SUPPLEMENTARY READING

Alexander Smith, *Inorganic Chemistry* (The Century Co., 1926), p. 193. (Here an account is given of the number of molecules in a liter.)

QUESTIONS

1. What volume does 50 liters of oxygen occupy at standard conditions?
2. Find the percentage composition of KNO_3 .
3. What per cent of Fe_2O_3 is iron?
4. What volume of hydrogen can be obtained from 100 grams of sulphuric acid, containing 50 per cent water?

5. What is the weight of 20 liters of CO_2 ?
6. A balloon has a capacity of 300,000 liters. What weight of hydrogen would fill it? What weight of water would give that much hydrogen?
7. How much pure limestone (CaCO_3) would have to be used to obtain 500 liters of CO_2 ?
8. The balloon mentioned in question 6 rises to a height where the atmospheric pressure is 400 mm. and the temperature — 20°C . What volume would the gas then occupy, if it were free?
9. An inner tube has a normal internal volume of 350 cu. in. How much air has been pumped in if the pressure gauge registers 50 lbs. (What is normal atmospheric pressure?)
10. 200 cc. of hydrogen at 18°C . and 745 mm. pressure would occupy what volume at 0°C . and 760 mm.?
11. What volume of chlorine, at normal temperature and pressure, would have to be forced into a tank to increase the pressure to 100 lbs. per square inch?
12. Calculate the volume of air needed to burn completely 1000 liters of gasoline vapor, assuming the latter to be C_7H_{16} .
13. A volume of gas measured 350 cc. when the barometer was 750 mm. and the thermometer, 25°C . What would be the volume at standard conditions?
14. Find the per cent of water in crystallized alum, 5 grams of which, on heating, finally weigh 2.57 g.
15. How much oxygen would unite with 150 g. of hydrogen?
16. A mixture of 75 cc. of oxygen and 25 cc. of hydrogen is exploded. What is left and how much?
17. What is the weight of air over one square inch at the sea level? What will the air weigh over a piece of ground $25' \times 100'$?
18. What volume of oxygen would be obtained from 100 grams of Na_2O_2 ?
19. What weight and volume of oxygen would completely burn up one kg. of coal containing 90 per cent C. and 2 per cent hydrogen?
20. How much copper will be extracted from one ton of pure CuS ? How much if the CuS were only 65 per cent pure?
21. How much water is present in one kg. of washing soda, $\text{Na}_2\text{CO}_3 \cdot 10 \text{H}_2\text{O}$?
22. Find the formula of the gas, one liter of which weighs 1.27 g., and whose composition is C, 92.3 per cent, H, 7.7 per cent.

ROBERT WAYNE EABY

2-1

GEORGE EDWARD MAYCOCK

CHAPTER XVI

SULPHUR. SULPHUR COMPOUNDS. SULPHURIC ACID

SULPHUR

184. This element finds itself in the same group in the periodic classification with oxygen; and while oxygen is a gas and sulphur is a solid, in chemical properties they have much in common. For example, metals such as copper and iron which combine with oxygen to form oxides also combine with sulphur to form sulphides :

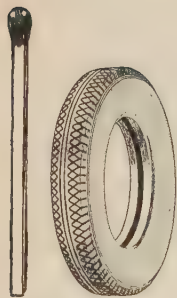
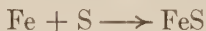
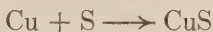


FIG. 166. Both of these contain sulphur.

Hydrogen combines with oxygen to form water and with sulphur to form hydrogen sulphide, H_2S . (At this stage compare also the properties of chlorine with those of oxygen and sulphur.)

Sulphur itself is an element of enormous importance. It is used in vast quantities in the manufacture of sulphuric acid, in the pulp and paper industry, in vulcanizing rubber, in destroying fungi, and in the making of matches and explosives. It is one of the essential elements necessary for plant growth. The skin contains appreciable quantities of this element, and there is good reason for supposing that it plays an important function in the activities of the cell.

185. Method of extraction. While deposits of sulphur are found in various volcanic regions throughout the world, such as those in Sicily and Japan, about 80 per cent of the world's supply is now

obtained from Louisiana and Texas, where, beneath layers of sand and clay, masses of sulphur are deposited. The sulphur is extracted by a process developed by *Frasch*, and since known by his name. The accompanying figure makes the essential points in the extraction clear. Hot water sent down through one pipe melts the sulphur, and compressed air sent down by another forces the melted sulphur to the top through a third pipe. The sulphur so obtained has less than 0.5 per cent impurity and is, as a rule, ready for the market.

186. Allotropic forms of sulphur.

Sulphur as ordinarily met with has a definite crystalline shape; it belongs to the *rhombic* variety (Fig. 168), and each crystal appears to consist of a double pyramid. This crystalline form is very sharply brought out by first dissolving the sulphur in some liquid in which the element is soluble (carbon disulphide is one of the best) and then allowing the solvent to evaporate.

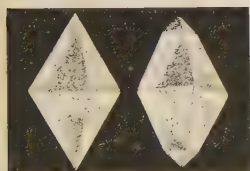


FIG. 168. Rhombic sulphur.

Quite a different crystalline variety of sulphur is obtained when the element is first melted and then allowed to cool; the crystals which look needle-shaped (Fig. 169) belong to the *prismatic* variety. Still a third form of sulphur, showing no crystalline characteristics, may be obtained by boiling the sulphur

and pouring the material into water; this gives us the *amorphous* variety (amorphous meaning "without definite crystalline structure").

The three forms of sulphur which have just been discussed —

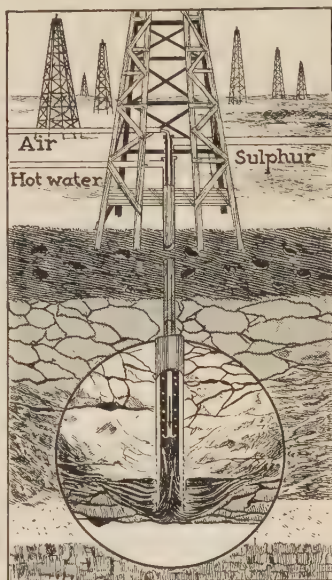


FIG. 167. Frasch process of extracting sulphur.

rhombic, prismatic, and amorphous — differ widely in physical properties (as in color and specific gravity, for example), yet in chemical properties they show little variation. For example, all three varieties unite with oxygen to form sulphur dioxide and with metals to form sulphides. Such varieties of one element, as sulphur shows, are spoken of as allotropic forms; and these different varieties are probably due to different arrangements of the atoms within the molecule, and to differences in the number of atoms in the molecule.



FIG. 169. Prismatic sulphur.

Sulphur may be bought in one of two varieties: either as *roll sulphur* or as *flowers of sulphur*, the latter consisting of very fine powder.

HYDROGEN SULPHIDE

187. We have said that sulphur and oxygen show striking similarities in chemical behavior. Hydrogen, we already know, combines with oxygen to form hydrogen oxide or water. We might therefore expect hydrogen to combine with sulphur to form a corresponding sulphur compound, *hydrogen sulphide*, H_2S ; and this is actually the case, though in practice it is much easier to cause hydrogen and oxygen to combine than to cause hydrogen and sulphur to combine. In any case, we use a far more convenient laboratory method for the preparation of hydrogen sulphide; it is by the action of hydrochloric or sulphuric acid on ferrous sulphide:

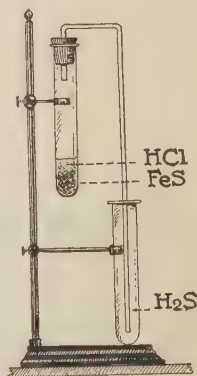
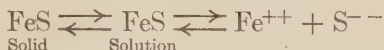


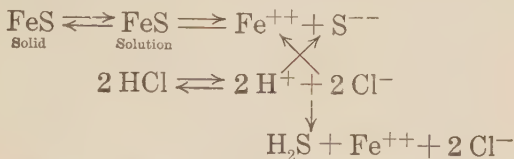
FIG. 170. Preparation of hydrogen sulphide.

The hydrogen sulphide — sometimes called sulphuretted hydrogen — being fairly soluble in water and heavier than air, is usually collected by downward displacement.

In terms of the ionic hypothesis, this reaction may be explained as follows: the ferrous sulphide, though called an insoluble substance, is slightly soluble, and the part which is soluble ionizes to a slight extent:



Now in the presence of hydrochloric acid, which readily ionizes ($\text{HCl} \longrightarrow \text{H}^+ + \text{Cl}^-$), hydrogen sulphide is formed; this, being very slightly ionized, is almost completely removed from the solution as un-ionized H_2S . Our law of mass action has already taught us that the removal of one of the ions formed speeds the reaction from left to right; so that more of the FeS ionizes. As fast as the S^{--} is removed to form H_2S , so fast does the FeS ionize. The reaction may then be represented:



188. Properties. Hydrogen sulphide is a far more poisonous gas than was supposed at one time; and the utmost precaution should be taken to prevent too much of it being absorbed. It has a very characteristic odor resembling that of rotten eggs (which, by the way, liberate hydrogen sulphide); and the odor is, as a rule, quite sufficient for identification purposes. It dissolves in water to some extent and forms a very weak acid, a slight dissociation taking place:



In its chemical behavior, its most marked property is that of a *reducing agent*. This is illustrated in a number of ways; for example, the gas will burn, combining, of course, with the oxygen of the air in doing so:



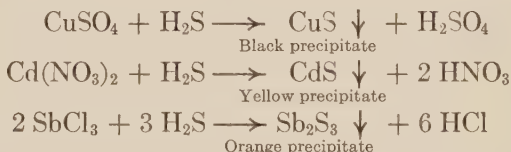
In the presence of a large excess of air, sulphur dioxide is formed in place of sulphur :



With hydrogen peroxide, H_2O_2 , which is a common oxidizing agent, the hydrogen sulphide reacts, as might be anticipated :



189. Uses. The value of hydrogen sulphide lies principally in its use as an important laboratory reagent in *identifying* many of the metals when these are combined in the form of compounds. To illustrate this, let us assume that we have three solutions: one copper sulphate, CuSO_4 , the next cadmium nitrate, $\text{Cd}(\text{NO}_3)_2$, and the third antimony chloride, SbCl_3 , and we wish to identify all three. By passing H_2S into each of these three solutions the following reactions take place :



The black precipitate, then, indicates the copper salt; the yellow precipitate the cadmium salt; and the orange precipitate the antimony salt.

Some natural waters, such as White Sulphur Springs, Saratoga, and Byron Springs, contain hydrogen sulphide.

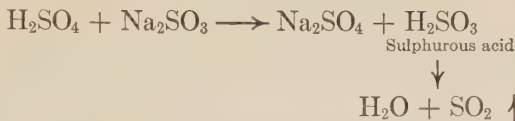
Many of the metals combine directly with sulphur and its compounds. The tarnishing of silver is due to the formation of black silver sulphide. Copper, lead, zinc, silver, mercury, and gold are extracted from sulphides found in the earth; in this extraction one of the by-products is sulphur dioxide.

SULPHUR DIOXIDE

190. Sulphur dioxide, SO_2 , can be prepared by the burning of sulphur in air or oxygen, as has already been suggested :



and this, indeed, is one of the commercial methods used in manufacturing the gas. For laboratory purposes, however, the gas is prepared in one of two ways: either by heating a sulphite with sulphuric acid:



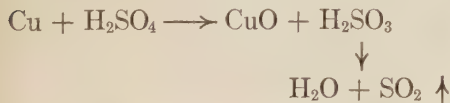
or by heating copper with concentrated sulphuric acid:



In either case the gas is collected by downward displacement (Fig. 171), since it is heavier than air and soluble in water.

When an acid acts on a sulphite, we suppose sulphurous acid, H_2SO_3 , to be one of the products; but since this acid is very slightly ionized and highly unstable, it readily breaks up into SO_2 and H_2O .

In the second method, where copper and concentrated sulphuric acid are used, the hot, concentrated acid acts like an oxidizing agent, oxidizing some of the copper to copper oxide and being itself reduced:



but a large part of the copper oxide so formed reacts with more sulphuric acid:



191. Properties. The gas has a very choking odor, like that obtained when matches containing sulphur are burned; and like chlorine and carbon dioxide can be comparatively easily liquefied —

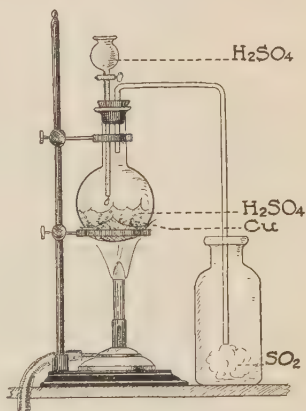
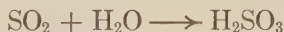


FIG. 171. Preparation of sulphur dioxide.

a fact made use of in a number of electrical household refrigerators, where sulphur dioxide takes the place of ammonia in ice making. When dissolved in water, it is assumed that a chemical reaction takes place :



and that sulphurous acid is formed — an acid which, by the way, has never been isolated. What is important, however, is that the solution turns blue litmus red and reacts with bases to form salts :



The calcium bisulphite formed is an important salt used in paper manufacture. Further, the solution of SO_2 in water shows very distinct reducing properties ; for like H_2S , also a reducing agent, it combines with oxygen, hydrogen peroxide, etc. :



192. Uses. The most important use of sulphur dioxide is in the manufacture of sulphuric acid, though large quantities are also used in the manufacture of paper and in bleaching wool and silk

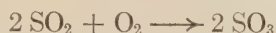
— materials which are too delicate for bleaching with chlorine. It is supposed that the sulphur dioxide (or, better still, the sulphurous acid) combines with the dye in the cloth to form a white product. Sulphur dioxide is both a disinfectant and a bleaching agent. It is used in preserving dried fruits, etc.



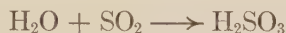
FIG. 172. Fruit is bleached by sulphur dioxide.

193. Besides sulphur dioxide there is another important oxide of sulphur, sulphur trioxide, SO_3 . This is formed in small quantities when sulphur burns in air or oxygen. (What is the main product ?) However, on a large scale the sulphur trioxide is made by passing

sulphur dioxide and air through a tube containing finely divided platinum (which acts as a catalytic agent) and which is heated to about $400^{\circ}\text{C}.$:



Now just as sulphur dioxide with water forms an acid :



so does sulphur trioxide with water form an acid :



only in the latter case we get sulphuric acid, the most important industrial product we have to deal with.

Sulphur dioxide and sulphur trioxide are called *acid anhydrides*, because they are substances which with water form acids.

SULPHURIC ACID

194. Sulphuric acid (the “oil of vitriol” of commerce) is actually made by a process which in principle has just been outlined. It is known as the *contact process* and is gradually superseding a much older method known as the *chamber process*, which, being rather involved, will not be discussed.

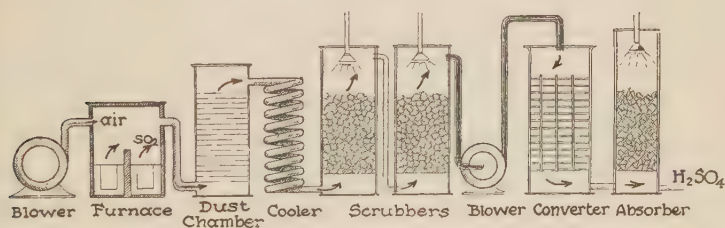


FIG. 173. The contact process for making sulphuric acid.

In actual practice, the sulphur is burned with an excess of air to form sulphur dioxide. Before converting the sulphur dioxide to sulphur trioxide, the gas has to be purified, otherwise the impurities attack the catalytic agent (platinum) and largely destroy its value. For this purpose the sulphur dioxide is first washed with water; then the latter is removed by passing the gas over concentrated

sulphuric acid (which acts as a dehydrating agent), and finally over ferric oxide, to remove arsenic compounds which, more than anything else, would quickly ruin the platinum if allowed to come in contact with it. The purified sulphur dioxide containing an excess of air is next passed through the tube containing the platinum, thereby forming sulphur trioxide.

The most obvious step now in the process would be to pass the sulphur trioxide directly into water and so form the sulphuric acid. Unfortunately, the sulphur trioxide is but slightly soluble in water; so that at best only a very dilute solution of sulphuric acid could be formed in this way. To overcome this handicap, the sulphur trioxide is first passed into concentrated sulphuric acid, in which it is made more soluble, forming *fuming* sulphuric acid:



This can now be mixed with water to form sulphuric acid:



If a relatively small quantity of water is added we get a concentrated sulphuric acid, and if a large quantity of water is added, we get a dilute acid. In other words, we can vary the concentration of the acid by varying the quantity of water added.

195. Properties. Among the common mineral acids, sulphuric is distinguished by the fact that it is a very heavy and a very oily liquid and that it generates considerable heat when mixed with water (giving rise, we now suppose, to some definite compound of the acid and water).

In dilute solutions it shows all the properties of a typical acid and reacts with metals above hydrogen (par. 44) to liberate hydrogen. In concentrated solutions, and when hot, it acts as a powerful oxidizing agent, reacting with copper (par. 190) and silver:



and with the halogen acids (pars. 134 and 137). The concentrated acid is also a very powerful *dehydrating agent*, which means that it absorbs water, or the elements forming water, from compounds.

We have seen how the sulphur dioxide is dried by passing it through concentrated sulphuric acid. Another example of this property is in the manufacture of explosives when the water which is formed as a by-product is removed by means of sulphuric acid. Its attraction for water is so great that a compound like ordinary sugar or paper, containing the elements carbon, hydrogen, and oxygen, will have its hydrogen and oxygen removed by sulphuric acid, leaving a deposit of carbon (black or charred deposit).

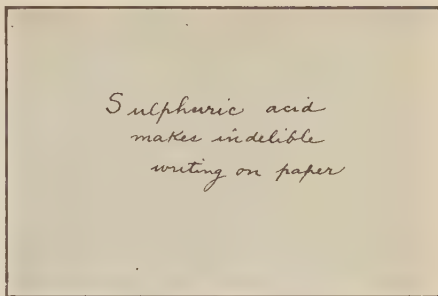


FIG. 174. Writing on paper with concentrated sulphuric acid, heating, and getting black writing.

196. Test for a sulphate. If we wish to identify sulphuric acid, we have to do two things: in the first place, we have to identify the hydrogen ion in the compound (which can be done by applying the usual tests for an acid); and in the second place, we have to test for the sulphate ion. This is done by adding barium chloride:



The white precipitate of barium sulphate so formed is insoluble in hydrochloric acid.

197. Uses. Sulphuric acid is more extensively used than any one single chemical. Because of its high boiling point, it is used in preparing nitric, hydrochloric, and other acids with lower boiling points. We use it to purify gasoline; to prepare superphosphate fertilizers; to prepare nitro compounds for explosives; to make various dyes; to clean metals; in storage batteries; etc.

QUESTIONS

1. In the Frasch method, what property of sulphur makes it possible to bring the element to the surface? Make your own diagram to explain how this works.

2. How could you change the amorphous form of sulphur into the prismatic and the latter into the rhombic?

3. How do we know that the various allotropic forms of sulphur are one and the same element?

4. By means of equations, show that sulphur and oxygen are similar in chemical behavior.

5. Using sulphuric acid and some sulphide other than iron, write an equation to represent the reaction of the two substances.

6. By means of equations, show that hydrogen and hydrogen sulphide are both reducing agents.

7. Remembering that hot, concentrated sulphuric acid is an oxidizing agent, what products might be formed when this reacts (a) with hydrogen sulphide; (b) with carbon?

8. The hydrogen prepared from sulphuric acid and metal sometimes smells like rotten eggs. To what is this due? Write an equation to represent the reaction.

9. Write equations to represent the reaction of hydrogen sulphide with HgCl_2 ; (b) BiCl_3 ; (c) K_2CO_3 . Look up the solubility table and indicate which products are soluble and which insoluble.

10. Write equations to represent the formation of sulphur dioxide from (a) sulphur; (b) hydrogen sulphide; (c) a sulphite; (d) sulphuric acid; (e) a metallic sulphide.

11. Write an equation to illustrate sulphur dioxide as a reducing agent. Do the same for carbon monoxide and carbon.

12. Potassium permanganate and nitric acid are oxidizing agents. What would be the effect of adding either to sulphurous acid?

13. Explain what happens when the following are passed into water: (a) H_2 ; (b) O_2 ; (c) H_2S ; (d) SO_2 ; (e) SO_3 ; (f) CO_2 ; (g) CO ; (h) N_2 ; (i) Cl_2 ; (j) Br_2 ; and (k) I_2 .

14. Explain why sulphuric acid is used in the manufacture of sulphuric acid.

15. If a beaker half full of concentrated sulphuric acid be exposed long enough, the beaker will be found to be full of liquid. Explain.

16. Desiccators are sometimes partially filled with sulphuric acid. Explain. (If necessary, consult the dictionary as to the meaning of the word "desiccator.")

17. Write equations for the reactions of (a) zinc; (b) iron; (c) magnesium; (d) mercury with first, dilute, and second, concentrated sulphuric acid.

18. Sulphuric acid (3 parts water to 1 part acid) used as ink makes a black writing. Explain.

19. Starting with sulphur, suggest methods for preparing (a) ferrous sulphide; (b) hydrogen sulphide; (c) sulphur dioxide; (d) sulphur trioxide; (e) sulphuric acid.

20. How would you test for the sulphate ion in a solution containing copper sulphate? Write the equation.
21. Suggest one reason why sulphuric acid can be so harmful and cause such pains when spilled upon the hands or face.
22. Compare the allotropic forms of sulphur with those of carbon.
23. In the preparation of H_2S and SO_2 , explain why the reactions go to an end.
24. Name three acids you have so far studied which are "weak acids." Explain what you mean by a "weak acid."
25. Give the physical properties of (a) hydrogen sulphide, (b) sulphur dioxide.
26. Find the percentage composition of sulphuric acid.
27. Two hundred cc. of sulphur dioxide weigh 0.575 g. Determine the molecular weight. Analysis shows that in sulphur dioxide there is 50 per cent sulphur and 50 per cent oxygen. Show how the formula is determined.
28. How many grams of hydrogen sulphide can be obtained from 10 grams of ferrous sulphide?
29. What volume of oxygen will be used up in the conversion of 100 liters of sulphur dioxide into sulphur trioxide? How much air would contain this volume of oxygen?

OPTIONAL QUESTIONS

30. How is a "blowout" in a rubber tire (shoe) repaired?
31. Chlorine is an oxidizing agent; sulphur dioxide is a reducing agent. Write reactions to represent the fundamental difference in chemical behavior when chlorine oxidizes and sulphur dioxide reduces.
32. In mixing sulphuric acid and water, the suggestion is given to pour sulphuric acid *into* water, rather than the reverse. Suggest the reason.
33. Selenium and tellurium belong to the same family in the periodic table that sulphur does. From your knowledge of the properties of sulphur, what properties of selenium and tellurium could you predict?
34. Refer to paragraph 187 and explain in a similar way (using ionic equations) the reaction between calcium carbonate and hydrochloric acid.
35. Suggest why sulphur dioxide, like carbon dioxide, is referred to as a "stable" substance.
36. By re-reading paragraph 197 and consulting the appropriate pages throughout the book [use the index], write a descriptive account of some of the important uses of sulphuric acid.
37. Write an equation to explain the formation of sulphur in sulphur springs.
38. How many kilograms of sulphuric acid can be obtained from the roasting of 100 kilograms of copper sulphide 60 per cent pure.

CHAPTER XVII

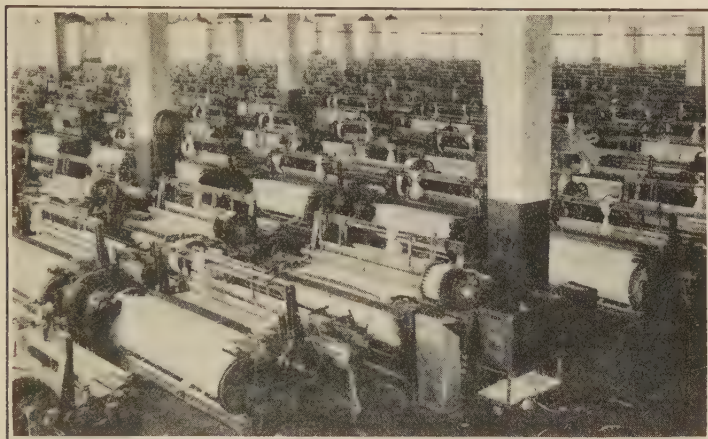
THE CLOTHES WE WEAR

198. Reasons for studying clothes. The choice of material for various articles of clothing is frequently determined by appearance alone; although, in most cases, compromise must be made with the pocketbook. Intelligent buying, however, ought certainly to take into consideration several other factors, chief among which are the hygienic nature of the clothing and its wearing qualities; for if the articles will last three times as long, it is cheap if its cost is only twice as much. Unquestionably, if we approach the problem of purchasing clothes with a greater knowledge of their properties, we shall be more successful, financially, and, what is of even more importance, we can protect and improve our health by proper choice of material to suit the conditions.

199. Materials used for clothing. Hundreds of different fabrics are used in the manufacture of clothing, but most of these are made of animal or plant gels of a fibrous nature, and those used are rather few in number. To be specific, most of our clothes are made of wool, cotton, linen, silk, leather, fur, and rubber. These are all plant or animal gels, and, except rubber, fibrous in structure. We shall now take up, one by one, the important materials used in the manufacture of clothing, and study their physical and chemical properties.

200. Cotton. Many plants scatter their seed by means of soft fluffy fibers, which, caught by the breeze, are carried aloft, and spread far and wide. Such are the so-called "money stealer", milkweed, and cotton plants. Before this seed dispersal takes place, however, the cotton bolls, as they are called, are picked, and

the cotton separated from the seed. From the latter, we get cottonseed oil, which is used for salad dressing and making fat for frying. The fiber is spun into threads and here the twists in the



Ewing Galloway

FIG. 175. Weaving.

fiber help, for the thread is held together only by friction; and with short fibers, like cotton, if the fiber is smooth and straight, there is less friction. Threads are woven into cloth, the many varieties of which differ chiefly in thickness of thread. For instance, muslin has a thin thread while canvas has a thick thread, but both are cotton.

201. Properties of cotton. Cotton fiber is white or slightly yellow, spiral shaped, soft, and silky in appearance, relatively short and weak. Chemically, it is a carbohydrate, almost pure cellulose, $C_6H_{10}O_5$, and rather inactive. At $250^\circ C$. it turns brown, is acted upon slowly by light, charred by sulphuric acid, unaffected by alkalis without air, but in the presence of air it is rapidly changed — a change which often plays havoc with our clothes when they are laundered. Cotton burns very readily, leav-



FIG. 176. Cotton fiber.

ing hardly any ash, a property which makes it rather dangerous for wearing apparel. In fact, there have been many cases of women and children being severely burned when their clothing caught fire. Cotton bleaches in chlorine, but the fiber is attacked and the life of the cloth is diminished. All laundries use chlorine solutions, which, although hygienic, shorten the life of the fabric.

202. Mercerized cotton. When cotton is stretched and dipped into strong caustic soda solution, the colloidal material tends to go into solution, strong alkalies favoring that action; and just as ordinary snow, when slightly melted by the sun and then frozen again, acquires a smooth surface, so the cotton fiber becomes smooth and shiny. It is now washed in dilute acid and water, to remove the alkali. This is mercerized cotton. It is stronger, and holds dye much better than cotton.

203. Other cellulose fibers. Kapok is used for pillows, mattresses, and life preservers. Jute, a coarse, strong East Indian fiber, is made into coarse bags, rugs, and twine. Manila, a Philippine fiber, is used for rope.

204. Linen. This is obtained from the stalks of flax. The fibers are longer and stronger than cotton, but harder to dye.



FIG. 177. Flax fiber.

Linen conducts heat better than any of the other fibers; therefore it is cool and suitable for summer clothing. It adsorbs water easily, a property which makes it desirable for towels. Because of its high luster, it is used to make table cloths.

205. Wool is obtained from sheep, goats, and camels. Under the microscope the fiber appears circular and curling, with overlapping scales. Each fiber acts like a steel spring. The length varies from one to eight inches, and it is fairly strong.

Wool is very hygroscopic; that is, it can adsorb as much as 50 per cent its weight in water without feeling moist. This, and its poor heat conductivity, make it feel dry and warm. Chemically, wool is

protein, of the same composition as hair, skin, and nail. It burns with difficulty, giving forth an unpleasant odor, and leaving much ash. To smother fire, therefore, never use a cotton article, but always wool, like a rug, a blanket, or an overcoat.

To test a piece of cloth for wool, it is first weighed and then boiled in a solution of sodium hydroxide. The wool dissolves completely, and if the cloth is now washed, dried, and weighed again, the loss in weight will equal the amount of wool. Since wool, therefore, dissolves in strong alkali, it becomes a problem to wash; for most cleansing agents, like borax, soda, lye, and soaps, are alkaline. Even boiling in plain water causes loss of

weight, called shrinking. To prevent this result, woolens should be washed in tepid water with good, non-alkaline soap.

Proteins are amphoteric, that is, they combine readily with acids as well as bases; and that accounts for the ease with which wool takes up many dyes. In

other words, wool is rather active chemically.

206. Quality of woolen fabrics. Wool seems to be a fiber very suitable for human apparel; but the quality of a fabric is not entirely measured by the percentage of wool. An article may be 100 per cent wool, and still it will not wear well. And why? If the fibers are very short, there is not much friction, and the threads pull apart. Wool is so much more expensive than cotton that old woolen clothing is sometimes torn apart into fibers and respun, by which process the fibers are shortened. Cloth made of this material is called shoddy, and while it has most



FIG. 178. A linen suit.

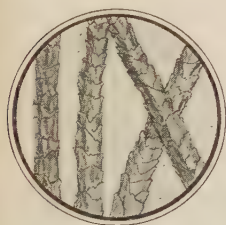


FIG. 179. Wool fiber.



FIG. 180.
Woolen clothing.

of the properties of wool, it is not strong. The strength may be increased by weaving cotton fiber with the wool.

207. Silk. Silk is also an animal product. It is a gelatinous secretion which is forced through small openings in the body of the silkworm and "sets" in air. If the silk gland of the silkworm is opened, the liquid can be removed and used to make fibers of greater thickness and strength. This substance is wound in the form of a cocoon about the caterpillar, serving as a protection during metamorphosis into a moth. Before the insect can emerge and spoil the cocoon, however, it is killed by heat. The cocoons are placed in hot water to soften the gum which holds the fibers, and then the double fiber is



FIG. 181. Sheep shearing.

unwound 300 to 1200 meters long. This is spun, and the raw silk marketed. Three thousand silkworms produce one pound of silk.

The fiber is smooth, lustrous, and of great strength; in fact, the strongest of all fibers. It burns very much like wool, with an unpleasant odor, and with a tiny black ball at the tip of the thread. It is a protein, turning yellow with nitric acid, soluble in concentrated hydrochloric and sulphuric acids. Crêpe is made by shrinking a silk goods in sulphuric acid. Silk is injured by salt. The latter, being present in perspiration, makes silk somewhat undesirable for underwear.

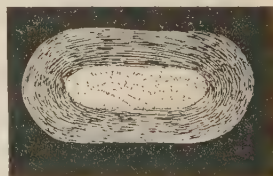


FIG. 182. Section of silk cocoon.

Weighting silk, that is, precipitating on it certain inorganic substances, like tin phosphate, gives to a thin, lifeless fabric, the body and feeling of heavy silk. This weakens the material, however. Such weighted silk can be distinguished from the pure article by burning, when it will not form the black ball; instead, it will glow like a metal wire, the large amount of ash being incombustible.

All silks, like wool, take dyes easily.

208. Rayon. The enormous demand for silk has stimulated the production of a great number of imitations. These may have the luster, but none have the strength. Rayon, however, is so much cheaper than natural silk that the manufacturer can afford to use far more material, and, therefore, produce a garment which is really stronger. Nearly all artificial silks are cellulose derivatives, resembling cotton in chemical properties. The cellulose is brought into the form of a gel as cellulose acetate or nitrate, squeezed through small openings to produce fibers, and then coagulated and spun into threads. (It is interesting to note that this process is the result of the observations of the process by which the silkworm produces the silk for its cocoon.) Artificial silks act like cotton; they burn easily, without ash, and do not take dyes readily. The chief commercial artificial silk is now known as rayon. The industry has grown from a production of a few thousand pounds in 1900 to 200,000,000 pounds in 1926.



FIG. 183. Necktie made of rayon.

209. Leather. In the earliest times, man made his clothing entirely of the skins of animals. Later, he learned how to cure the hides, to prevent decay, and also to toughen the leather. This process is called tanning. The hide is first soaked in slaked lime (an alkali) until it swells and softens (starting colloidal solution), when the hair can easily be removed. This is called "depilation." The lime is then neutralized by weak acids, and the hide is soaked in the tannin (usually obtained from oak or hemlock bark) for about two months. The gelatinous material of the hide is coagulated, which makes it hard and insoluble, and leather results. Sole leather is given a specially long tanning, until, on boiling a sample, no gelatin can be extracted. Chrome tanning is done with hydrochloric acid and potassium dichromate ($K_2Cr_2O_7$) in place of vegetable tans. Alum and even ferric salts are sometimes used.

(Notice that in each case the elements Cr, Al, Fe, all with a valence of three, are effective coagulating agents, and recall that the precipitating power of an ion increases with the valence.) The dark green color of the chrome tanned leather is due to chromium hydroxide, $\text{Cr}(\text{OH})_3$, deposited in the fibers.

210. Some common leathers. *Russia calf* is tanned with willow bark. *Cordovan* is horsehide. *Kid* is goat or sheep skin treated



Ewing Galloway

FIG. 184. Hides hung up before tanning.

with sour bran to remove lime, and then tanned with alum, salt, flour, and egg yolk. *Suède* is made from young calves' hides, but finished on the inside instead of the hair side. *Patent leather* is made by treating certain kinds of leather with coatings of varnish. *Morocco* is made from goat skins.

Furs are hides that are tanned without removing the hair. An alum tannage is used, and this dissolves out when the fur is drenched, causing the leather to set hard.

211. Rubber clothing. We shall not at this stage take up the chemistry of rubber, but we ought here to refer to the use of rubber in articles of clothing: waterproof cloth for rubber coats and shoes. Such articles are made by spreading on the cloth a layer of a mixture of rubber and sulphur, which is then vulcanized. Rubber is soluble in gasoline, ether, carbon tetrachloride, and many other organic liquids, but it is insoluble in alcohol. Such solvents should, therefore, not be brought into contact with rubber articles; though the knowledge is useful in mending such fabrics.

212. Fireproofing. Fabrics may be fireproofed (that is, they will not flame up, like cotton, although they will smolder) by being dipped in a solution of ammonium phosphate, alum, or borax. Ammonium phosphate is volatile, and this vapor surrounds the fabric, keeping off oxygen. But these salts are soluble; and a more effective but more expensive fireproofing consists in precipitating on the fibers tin oxide, which is insoluble.

213. Sizing. Various materials are put on cloth to give it body or stiffness. Among these are starch, gum, and paraffin. This is called sizing. Magnesium chloride is sometimes added, which, being deliquescent, takes up moisture and makes the fibers interlock.

214. Dyeing. Silk and wool are chemically active, also they are more nearly colloidal and adsorb better than cotton and linen. Depending on these facts and on the nature of dyestuffs, there are three systems of dyeing:

1. *Direct dyeing.* Most dyes are colloids which can be precipitated or "salted out" of solution by addition of a salt; and for silk and wool, which adsorb well, direct dyeing is perfectly possible. There are a few dyes, like congo red, which are direct for cotton.

2. *Insolubility.* If the insides of the fibers can be filled with a solution which can then be precipitated as an insoluble color, a fast dye will result. This is done with indigo.

3. *Dyeing with a mordant.* Whenever the fabric has no adsorptive power for the dye, another colloid, itself of great adsorptive

power, called a *mordant*, is used to hold the dye. The mixture of mordant and adsorbed dye is called a *lake*. Aluminum hydroxide and tannic acid are two such mordants. In most cases cotton fabrics require mordanting. Printing on cloth is done by stamping the design on one side with the mordant and then dyeing. The cloth does not hold the color where no mordant is attached.

215. Fastness of color. No dye is unchanging, but if it lasts as long as the fabric, it is certainly satisfactory. The action of any chemical reagent may affect dyes; colored fabrics, therefore, should not be treated with alkali or acid. Mild soaps should be used, and, after removing the soap, the dye may often be coagulated by rinsing in salt water. This may prevent "running" of the color, or "bleeding." If a dye is discolored by the action of an acid, such as fruit juices or vinegar, the original color may often be restored with a few drops of alkali, like ammonium hydroxide or sodium bicarbonate; and vice versa, if alkali discolors, use vinegar or lemon juice to restore the color.

216. Simple tests for fabrics. When one enters a store to buy clothing, many of the tests described cannot be used for lack of a chemical laboratory. This paragraph, therefore, will give methods of testing fabrics with available materials.

Cotton. If a thread is separated into its fiber by untwisting, each fiber waves about gently. When pushed down it does not spring back, but returns slowly. Set afire, it burns quickly and completely, with almost no smell or ash.

Linen. It is practically impossible to distinguish linen from cotton except by a microscopic test. Some depend on the relative unevenness of linen threads to distinguish them from cotton; while others use the wet finger test, depending on the fact that the water will penetrate linen faster than cotton. But these tests are unreliable.

Wool. If a thread of wool is untwisted and each fiber held up, it can readily be distinguished; it is kinky, and when pulled down, springs back like steel. On burning, it produces a bad odor and

does not continue to burn, but forms a black ball at the end. With a little practice, it is only necessary to look at the fibers to distinguish each one from cotton.

217. How to tell the per cent of wool in a fabric. Let us suppose that one wishes to test a piece of cloth for the per cent of wool as well as length of fiber. Secure the longest thread possible, and untwist it, so as not to tear the fibers. Hold up each fiber in turn, and inspect it. In this way count the number of fibers which are wool and the number which are cotton. Then select a thread from those running at right angles to the first, and repeat. The results will tell the length of fiber and the number of wool and cotton fibers, from which the percentage of wool can be calculated, approximately.

218. How to tell the quality of a silk fabric. Here the problem is to distinguish pure silk from loaded silk and from artificial silk. The pure silk fiber burns with a black ball, whereas the artificial silk burns up entirely, like cotton; and the loaded silk glows like a red hot wire, but does not seem to burn. A lighted match, therefore, enables one to tell, easily and quickly, these three kinds of silk.

219. Removal of stains. Stains are easily removed from woollens, for when the proper solvent is found, the great surface of the material adsorbs, and so spreads, the solution uniformly. For instance, to remove a grease spot from a piece of woolen goods, it is only necessary to apply a little gasoline or carbon tetrachloride, and rub gently. The solution of grease sinks in and covers a large amount of wool fiber, and when the solvent evaporates, it leaves a very thin layer of grease on a large surface of wool; so it is really not removed, but imperceptible. Whereas, with a thin silk goods, the effect would be simply to enlarge the spot. With such an article, dry cleaning is the only perfect treatment, for then the stain is dissolved in a large volume of solvent, the goods wrung out gently, and the rest of the solvent evaporated. Of course some of the grease is deposited on the fabric by evaporation of the solvent;

but in the first place, the amount is small, since the volume of solvent is very great; and secondly, it is uniformly spread, so that there is now no spot. Spots on silks may be removed fairly well without dry cleaning, however. The spot is moistened with the solvent and rubbed with a dry cloth until the solvent has evaporated. This process may have to be repeated to prevent a ring.

If the color has been removed, the looks of the article may be improved with a little vinegar; or better yet, the article may be retouched with a solution of some colored substance in benzene.

If the cause of the spot is known, it is usually easy to remove. If unknown, try first water, — except on silks, — then carbon tetrachloride. On the opposite page is given a table which will enable one to remove most stains from various fabrics.

220. Washing clothes. From the earliest times, the necessity of washing has been recognized. Before the invention of soap, washing was merely a mechanical process, the clothes being immersed in water, usually in a brook, and rubbed on stones. The modern washboard imitates the action of the stones, but we have added soap and other substances. The action of soap in cleaning has been explained as due to its emulsifying action. Dirt and grease are insoluble in water, and therefore tend to cling to fabrics; but grease is soluble in alkaline solutions (like soap or soda), while dirt particles are held in suspension in soap water. Mechanical action (rubbing) assists in loosening and breaking up dirt, so that it is more readily emulsified. Boiling also assists, by increasing the solubility as well as by increasing the movement of the material.

221. Action of washing on fabrics. Cotton and linen, being inactive chemically, may be treated quite roughly, using almost any grade of soap, or even soda. Wool must be handled with great care. It must be washed only with neutral soap, never soda, and not boiled, since wool combines chemically with all alkalis, especially at higher temperatures, causing loss in weight ("shrinking"). Silk is very much like wool in this respect.

HOW TO REMOVE STAINS

STAIN	COTTON OR LINEN	WOOL	SILK
Blood	Soap and water; Javelle water, NaOCl, if necessary	Soap and water	Soap and water, then alcohol
Chewing gum	Toluene or ether	Same as for cotton	Same as for cotton
Chocolate	Boiling water and NaOCl; or glycerine, then water	Glycerine, then water	Same as for wool
Coffee	Same as for chocolate	Same as for chocolate	Same as for chocolate
Fruit	If fresh, water; or lemon juice, vinegar, or oxalic acid	Same as for cotton	Solution of soap in alcohol
Grass	Soap in alcohol	Same as for cotton	Same as for cotton
Grease	carbon tetrachloride, CCl ₄	CCl ₄	CCl ₄
Ice cream soda	Water	Water	Water
Ink	When fresh, soap and water; otherwise, NaOCl, then oxalic acid, then water. Repeat	H ₂ O ₂ , then oxalic acid	If fresh, soap and water; if not, same as for wool
Paint	Soap and water, when fresh CCl ₄	Same as for cotton	Same as for cotton
Rubber (adhesive plaster)	CCl ₄	CCl ₄	CCl ₄
Rust	Lemon, vinegar, or oxalic acid	Same as for cotton	Same as for cotton
Scorch	H ₂ O ₂ or NaOCl	H ₂ O ₂	H ₂ O ₂
Syrup	Water	Water	Water
Tar	CCl ₄	CCl ₄	CCl ₄
Tea	If fresh, water. Otherwise same as for chocolate	Same as for cotton	Same as for cotton
Varnish	Alcohol or CCl ₄	Same as for cotton	Same as cotton

222. Other cleansing agents. Many other materials are used in cleaning clothes, but most of them are alkaline. Lye is the strongest of these. A very small amount may be used with cotton fabrics, but it is a powerful caustic, injuring the skin, and its use is not recommended. Washing soda is next in power. It is much stronger than soap in alkalinity, and is used in most washing powders. It is perfectly safe for cotton and linen, but not for silk or wool. Ammonia, as it is called (strictly ammonium hydroxide), is excellent for cottons. Being volatile, any excess used evaporates, and does not remain to injure the fabric. Borax is perhaps the weakest of the alkalies used in washing. All alkalies have the advantage of softening hard water so that their use is advisable for most localities.

223. Bleaching. Javelle water, NaClO , bleaches because of its chlorine. Most laundries use chlorine, in some form, to whiten the clothes, and this would not be objectionable (especially since chlorine is a powerful germicide) if it did not also weaken the fabric. Used with care, no great injury is done; but a careless laundryman, using too much, may take most of the life out of a batch of clothes. Laundries buy their chlorine as bleaching powder (CaOCl_2), javelle water (NaClO), or liquefied chlorine gas; but some are now installing vats in which the bleaching is done at the anode in an electrolyte of sodium chloride, chlorine being liberated only when needed, by turning on the current.

SUPPLEMENTARY READING

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- H. E. Howe, *Chemistry in Industry* (Chemical Foundation, 1924): Leather, chap. xii; Textiles, chap. xxi; Rayon, chap. xviii.
- Kinne and Cooley, *Shelter and Clothing* (The Macmillan Co., 1923).
- H. E. Howe, *Chemistry in the World's Work* (D. Van Nostrand Co., 1926), chap. v.
- E. E. Slosson, *Creative Chemistry* (Century Co., 1919), chap. vi (Cellulose).

- ¹⁶ "Materials for the Household," *Circular of the Bureau of Standards*, No. 70 (Government Printing Office, 1917): pp. 140-145 (rubber); pp. 145-151 (leather); pp. 151-163 (textiles); pp. 192-197 (cleansing agents).
- Garard, *Applied Chemistry* (The Macmillan Co., 1924): pp. 273-291 (cleansing agents); pp. 291-320 (textiles and dyeing); pp. 320-334 (leather and tanning).
- Svante Arrhenius, *Chemistry in Modern Life* (D. Van Nostrand Co., 1925), chap. xiii (Cellulose and rubber).
- F. L. Seymour-Jones, "The Beginning of Leather Chemistry," *Journal of Chemical Education*, (1927) vol. IV, p. 831.

QUESTIONS

1. What qualities of clothing must be considered — besides price — to make a good selection?
2. Name five natural substances used for clothing.
3. In tabular form, name ten woven fabrics, with the name of the natural material from which each is made. If made of the same material, explain the difference.
4. What is mercerized cotton? How could you identify it? How does it compare with cotton and silk in cost, wearing qualities, and appearance?
5. What danger arises from wearing loose cotton garments near a cook stove?
6. If a person's clothing caught fire, what kind of cloth should be used to smother it? Why?
7. Why is linen used for towels?
8. What fabric is best for winter underclothing?
9. What fabric is best for summer underclothing?
10. Why does a cotton undergarment feel wet and sticky in hot weather? What kind would feel dry?
11. Describe how a garment can be tested, without using chemicals, for its approximate per cent of wool.
12. Why do cotton fabrics fade easily?
13. What is shoddy? Is it warm? Why? Why does it not wear well?
14. Which fiber is strongest? Account for the poor wearing qualities of most articles of clothing made of this fiber.
15. How would you assure yourself that a necktie was silk?
16. Describe how to test for loaded silk.
17. How can rayon be distinguished from pure silk?
18. What is accomplished by tanning hides?
19. Why is lime used to remove hair from hides?

20. What is the difference between fur and leather?
21. Explain why cotton needs a mordant for dyeing.
22. Why does the attempt to remove a grease spot often leave a ring?
23. For what fabrics is dry cleaning advisable? Explain.
24. Why is a stain easily removed from a woolen suit?
25. Tell how to remove chewing gum from the hair.
26. Name some common cleaning fluids. Which one is dangerous? Explain. Which one is safe? Why?
27. Describe how to remove a rust stain from a linen handkerchief.
28. How would you remove a drop of paraffin from a dress?
29. What causes shrinking?
30. Explain the action of soap in washing.
31. In what way is "ammonia" better than washing soda for cleaning?
32. Give the advantage and disadvantage of the use of chlorine by laundries. How may the disadvantage be minimized?
33. What is meant by a fast color?
34. Explain why sodium hydroxide is more strongly alkaline than washing soda.
35. Cellulose is essentially $C_6H_{10}O_5$. What are the products of combustion?

CHAPTER XVIII

NITROGEN. OXIDES OF NITROGEN. NITRIC ACID. AMMONIA

NITROGEN

224. We have already referred to nitrogen in connection with our discussion of the atmosphere (chap. iv) as an element present in great abundance in the air, and we have also had occasion to refer to its inertness or its reluctance to combine with other elements. It is, we have pointed out, a non-supporter of combustion. It is taken in by us and by all things that have life, together with the oxygen, but is not used. Nitrogen can, under certain conditions, be made to combine with other elements. For example, during a thunderstorm, whenever there are flashes of lightning, a small portion of the oxygen and nitrogen do actually combine with one another and this is the starting point of an important commercial process.

225. Argon. Nitrogen can also be made to combine with a number of the metals to form nitrides. Magnesium, for example, forms magnesium nitride, Mg_3N_2 . This compound is of historic interest for this reason: When Ramsay was studying the gases of the atmosphere, he first passed the air over heated copper and thereby made the oxygen and copper combine, forming copper oxide. The residual gas, which was supposed to be nitrogen, was next repeatedly passed over heated magnesium. The nitrogen combined with the magnesium to form magnesium nitride. Ramsay, however, found that a small portion of what was believed to be nitrogen refused to combine with the magnesium. This small portion turned out to be quite another element — argon.

226. Nitrogen and life. Nitrogen is often spoken of as an indispensable element of living tissues. As a matter of fact, neither the plant nor the animal is able to absorb and utilize the element as such. But in the form of nitrates or ammonium salts, nitrogen is indispensable to plants; and in the form of protein, the nitrogen is equally indispensable to animals.

227. Properties and uses. Since both the oxygen and the nitrogen constitute the bulk of the atmosphere, it can be foretold that the nitrogen, like oxygen, is a colorless gas; and since so large a part of the atmosphere covers the oceans and seas, and yet nearly 80 per cent of this atmosphere is nitrogen, one can also draw the conclusion that nitrogen, like oxygen, is not very soluble in water.

Enormous quantities of nitrogen are required in the preparation of ammonia and in the preparation of cyanamide. The gas is also used to fill certain types of electric lamps.

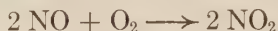
OXIDES OF NITROGEN

228. Nitrous oxide, N_2O , also called "laughing gas," is an important anæsthetic the properties of which were first described by Sir Humphry Davy. It is made by heating ammonium nitrate:



The gas resembles oxygen in supporting combustion.

229. Nitric oxide, NO , is commonly prepared by the action of copper on nitric acid. Though colorless, it becomes brown in color when exposed to the air. This is due to the fact that the nitric oxide combines with the oxygen to form *nitrogen peroxide*:



and it is this gas which is brown in color and which has a pungent odor. The nitrogen peroxide, unlike the nitric oxide, is soluble in water, forming nitric and nitrous acids:



During thunderstorms, nitrogen peroxide is formed and is dissolved by the rain, which explains the small quantities of these acids present in rain water.

230. Law of multiple proportions. We have just seen that nitrogen combines with oxygen to form a number of different compounds — N_2O , NO , and NO_2 . Two others are also known: nitrogen trioxide, N_2O_3 , and nitrogen pentoxide, N_2O_5 . If, for the sake of convenience in comparing these substances, we double the atoms in NO and NO_2 , giving us N_2O_2 and N_2O_4 respectively, then we have the list

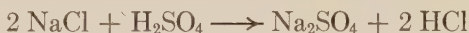
	N	O
Nitrous oxide	28	16
Nitric oxide	28	32
Nitrogen trioxide	28	48
Nitrogen peroxide	28	64
Nitrogen pentoxide	28	80

the numbers representing the weight relationships. For example, in nitrous oxide, we can say that for every 28 parts by weight of nitrogen there are 16 parts by weight of oxygen. Similarly with the other compounds. Now you will notice that with the weight of one of these elements fixed — in this case, the nitrogen at 28 — the numbers for the oxygen in these compounds are in the ratio of small whole numbers; for 16 : 32 : 48 : 64 : 80 is in the relationship of 1 : 2 : 3 : 4 : 5. This illustrates the *law of multiple proportions*, which may now be restated as follows: when two elements *A* and *B* combine in more than one proportion (to form more than one compound), if we consider a fixed weight for *A*, then the weights of *B* which unite with this fixed weight of *A* are in the ratio of small whole numbers.

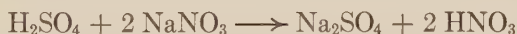
NITRIC ACID

231. Nitric acid. HNO_3 is one of the most important of the common acids. The student recognizes it quite early in his course, because if he gets it on his skin it leaves a yellow mark.

232. Preparation. Nitric acid is prepared commercially in a number of ways. One way is similar to that used for the preparation of hydrochloric acid. It will be recalled that in the preparation of this latter acid, sodium chloride and concentrated sulphuric acid are heated :



Likewise, if in the place of sodium chloride we use sodium nitrate (the Chile salt-peter found in such abundance in Chile), we get nitric acid :



In reality we get sodium acid sulphate, NaHSO_4 , but for the sake of simplicity we use the equation given. This method serves both commercially and as a laboratory method (Fig. 185).

Another commercial method, and one introduced more recently, is to oxidize ammonia in the presence of a catalytic agent :



This process, suggested by Ostwald, was developed in Germany on a large scale during the late war. Platinum was the catalytic

agent found most suitable. In the actual operation, the ammonia and oxygen (or air) are passed through a tube containing finely divided platinum heated to about 700°C . The oxides of nitrogen which are formed are dissolved in water to form nitric acid.

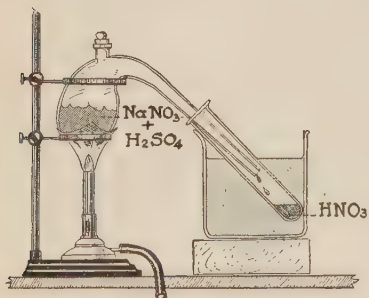


FIG. 185. Preparation of nitric acid.

According to Dr. C. L. Parsons, this process has been developed to such a degree, that nitric acid can now be manufactured from ammonia for \$30 less per ton than it can be made from Chilean nitrate (which sells around 2¢ a pound). Nitrogen in the form of

ammonia is now obtainable for less than half the price of nitrogen in the form of Chilean nitrate. All of which points to the probability that, sooner or later, this process, the Ostwald process for manufacturing nitric acid, will entirely supplant the much older process of extracting the acid from sodium nitrate.

Still another method, one by which the nitric acid is obtained from the atmosphere, will be discussed later.

233. Properties. Although nitric acid is frequently seen as a yellow liquid, it is quite colorless when pure, just like hydrochloric and sulphuric acids. However, on standing, slow decomposition occurs, forming, among other things, nitrogen peroxide, which, as we have already seen, is yellow in color. It is this nitrogen peroxide dissolved in nitric acid which gives the latter its yellow color.

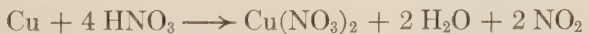
Chemically, nitric acid is a most active oxidizing agent. This is best illustrated in its action on metals. When dilute hydrochloric or dilute sulphuric acid reacts with zinc, hydrogen is evolved; but when dilute nitric acid and zinc are brought together, the hydrogen first evolved is immediately oxidized to water. Even metals below hydrogen in the series (par. 44), such as silver and copper, are attacked by nitric acid:



(The solution of copper nitrate is greenish blue in color.)



Concentrated acid on copper yields a slightly different result:



Platinum and gold, however, are not attacked except by *aqua regia*, which is a mixture of nitric and hydrochloric acids and may be roughly represented as:



The gold and the platinum are here attacked by the chlorine just as soon as it is evolved in the reaction (the chlorine being in the "atomic" state); and the corresponding chlorides are formed.

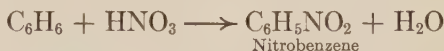
Brass, which is often used as an imitation for gold, may be distinguished from the latter by adding nitric acid, the brass turning green.

234. Test for a nitrate. The NO_3^- in nitric acid or in any soluble nitrate may be tested for by the addition of ferrous sulphate and concentrated sulphuric acid, when a brown ring is formed between the H_2SO_4 and the watery layer.



FIG. 186. Test for a nitrate.

235. Uses of nitric acid. One very important use is in the manufacture of nitro compounds. For example, benzene when treated with nitric acid forms nitrobenzene ;



Toluene treated with nitric acid under the proper conditions forms trinitro-toluene, or T.N.T., a powerful explosive. Similarly, glycerol and cellulose form nitro-glycerine and nitro-cellulose. These nitro compounds and nitrates are important, therefore, because many of them are powerful explosives.

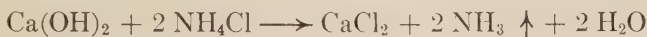
But nitric acid finds wide use in other industries, such as the manufacture of dyes, films, rayon, celluloid, collodion, etc.

The vigorous action of nitric acid on metals finds an important application in the process known as "etching." Here a plate of copper or zinc is covered with a thin layer of paraffin or asphalt and the outline then made on the paraffin, using a sharp instrument, and thereby exposing parts of the metal. The plate is next covered with nitric acid, which attacks the exposed parts of the metal. This plate may then be used to print a large number of copies.

AMMONIA

236. Ammonia, now so extensively used in refrigeration and with water as a cleansing agent, is formed whenever organic matter decays. It is prepared in very large quantities in the

course of the distillation of coal. It is also prepared in enormous quantities by the Haber process, in which nitrogen and hydrogen are directly combined; and by the decomposition of calcium cyanamide with water. In the laboratory it is most conveniently prepared by heating an ammonium salt with a non-volatile base (Fig. 187):



This reaction forms the basis for the *test for any ammonium salt*: heat the ammonium salt with a base and ammonia will be evolved, which can be recognized by its pungent odor, by the fact that it turns red litmus blue when the ammonia is dissolved in water, or by the dense white fumes formed when the ammonia comes in contact with hydrochloric acid.

237. Properties. Ammonia is a colorless gas with a pungent odor. Its density is less than that of air and it is therefore collected by upward displacement (see Fig. 187). It is extremely soluble in water, forming *ammonium hydroxide* (household ammonia):



which is a typical base, ionizing to some extent thus:



NH_3 is therefore a *basic anhydride*.

238. Artificial ice. The use of ammonia in refrigeration depends upon the relative ease with which ammonia gas can be converted into ammonia liquid by pressure and cooling. The principle of the procedure can be illustrated with the help of the diagram on the next page:

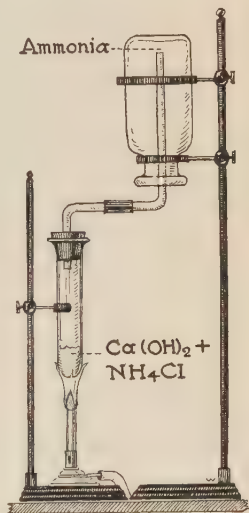


FIG. 187. Preparation of ammonia.

Liquid ammonia is first prepared by compressing and cooling the gas and forcing the liquid so obtained through a narrow tube which ends in a wider tube where, having room for expansion, the liquid evaporates and becomes gas once more. But in doing

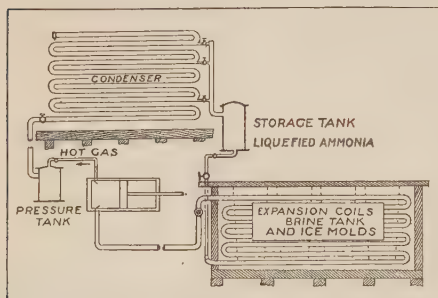


FIG. 188. Manufacture of artificial ice.

so it absorbs heat from the surroundings. If, therefore, the larger tube is surrounded by water, the ammonia will obtain its heat from the water; which means that the water will become cooler and finally turn into ice.

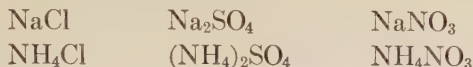
In actual practice the ammonia under pressure is allowed to expand when sent through the coil. The coil is surrounded by a solution of salt, the temperature of which is considerably reduced, so that when cans of water are placed in the salt solution (brine) their contents freeze to ice. This ice, which, as you may notice, never comes in contact with the ammonia, is what the housewife and others use. In cold storage places brine is pumped through pipes in the rooms which are to be cooled.

In addition to the use of ammonia in refrigeration, it is also used in the manufacture of sodium carbonate and sodium bicarbonate, and, as we have already seen, in the manufacture of nitric acid.

239. Among the salts of ammonia, *ammonium sulphate*, $(\text{NH}_4)_2\text{SO}_4$, is used as a fertilizer; *ammonium chloride* (sal ammoniac), NH_4Cl , is used in dry cells and in soldering; and *ammonium nitrate*, NH_4NO_3 , was used in the war for explosive mixtures.

The word "ammonia" stands for NH_3 , a gas. What is often called "ammonia" is really the solution of ammonia in water;

that is, ammonium hydroxide, NH_4OH . "Ammonium," NH_4 , really acts like a metal; for example,



Here, as can be seen, the NH_4 takes the place of the Na, a metal.

240. The ammonium group, NH_4 , has never been isolated. It is known only in the form of its compounds, some of which we have just mentioned. In its chemical reactions it behaves as if it were an element, the N and the H never parting company. In this respect the ammonium group is similar to other groups which have already been referred to, such as CO_3 , SO_4 , NO_3 , etc.

(Recently an entirely new chemistry has arisen by the substitution of liquid ammonia in the place of water as a solvent. In our ordinary experiments we allow metals, acids, bases, and salts to react with water. What will happen if such substances

are allowed to react in a medium of liquid ammonia instead of water? Some very interesting results have already been obtained. The student is referred to the supplementary reading at the end of this chapter.)

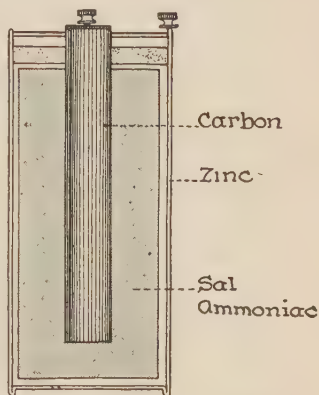


FIG. 189. A dry cell containing sal ammoniac.

SUPPLEMENTARY READING

- R. H. Bullard, "Liquid Ammonia as a Solvent," *Journal of Chemical Education*, October, 1925, p. 924.
 C. L. Parsons, "Nitric Acid from Ammonia," *Industrial and Engineering Chemistry*, vol. XIX (1927), p. 789.
 F. G. Keyes, "Chemistry in Refrigeration," *Chemistry and Industry*, vol. II, p. 330.

QUESTIONS

1. Illustrate by example the "inertness" of nitrogen, and show how the gas may be obtained from the air.
2. If we inhale a liter of nitrogen, how much of the gas, if any, will we exhale?
3. Describe some of the uses of nitrogen.
4. Give the equations for the preparation of nitric oxide and nitrogen peroxide.
5. Give, with equations, two industrial methods used for the preparation of nitric acid.
6. Illustrate by examples the actions of nitric acid as an oxidizing agent.
7. Give the test for a nitrate.
8. Give a method used for preparing ammonia (*a*) from its elements, (*b*) from an ammonium compound. Show how you can identify any ammonium salt.
9. Describe the method employed in the preparation of liquid ammonia.
10. Complete the following equations :
 - (*a*) Dilute nitric acid plus copper —.
 - (*b*) Dilute nitric acid plus silver —.
 - (*c*) Dilute nitric acid plus potassium hydroxide —.
 - (*d*) Nitric acid plus hydrochloric acid —.
11. How could you detect the brass in a gold-plated watch case?
12. Write the equations for (*a*) the laboratory preparation of ammonia, (*b*) the action of ammonia gas on hydrogen chloride, (*c*) the reaction between hydrogen sulphide and nitric acid.
13. Mention some important uses of ammonia gas ; of ammonium compounds.
14. Account for the instability of nitrogen compounds. Contrast this with the stability of hydrogen fluoride and explain.
15. Write equations to represent the reaction between (*a*) sulphuric acid and charcoal, (*b*) hydrogen peroxide and hydrogen sulphide, and (*c*) nitric acid and hydrogen sulphide.
16. Given bottles containing three different acids, — nitric, hydrochloric, and sulphuric, — show how you would identify each of these (*a*) by physical tests, (*b*) by chemical tests.
17. Write ionic equations to represent the preparation of (*a*) nitric acid and (*b*) ammonia. Why do these reactions go to an end?
18. Show by means of equations how the following salts can be prepared: ammonium nitrate, ammonium sulphate, and ammonium chloride.

19. Write equations to show that ammonia, sulphur dioxide, and carbon dioxide are anhydrides.

20. How many kilograms of Chile saltpeter must be used to make 100 kilograms of nitric acid?

21. How many liters of ammonia gas could be obtained from 20 grams of ammonium sulphate? How much does this ammonia weigh?

22. Using the Haber process, how many liters of ammonia will be obtained from 500 liters of hydrogen?

23. What is the percentage of oxygen in nitric acid?

OPTIONAL QUESTIONS

24. How was argon isolated?

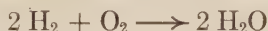
25. Using the two oxides of carbon, explain the law of multiple proportions.

26. Write an essay dealing with nitric acid in its relation to explosives.

CHAPTER XIX

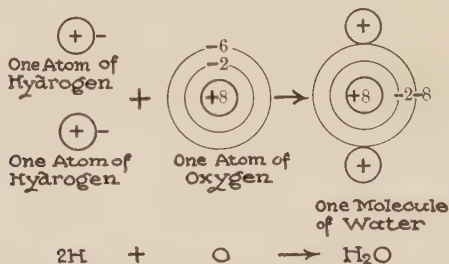
OXIDATION AND REDUCTION

241. The oxidation of hydrogen. Consider the equation :



In terms of our older definition, we would say that hydrogen is oxidized to water. Since in any chemical reaction every oxidation is accompanied by a corresponding reduction (par. 45), we may say that oxygen, in turn, is reduced to water. Hydrogen, we have learned, is a reducing agent ; oxygen, an oxidizing agent.

Let us consider the combination of hydrogen and oxygen to form water in terms of our electron conception :



It will be noticed that by bringing the two electrons from the two hydrogen atoms closer to the six electrons from the oxygen atom, we approach a stable configuration. The H atom is now H^+ , the O atom is now O^{--} . The H^+ and O^{--} are really *ions* in structure ; but, unlike the ions in solution, they are held together by electrical forces. For the sake of simplifying the balancing of equations by the electron method, we will write the equation, remembering that elements as such have zero valence (due to an equal

number of electrons and protons), and that their valence becomes one or more when they become part of a compound :



In this equation the 2H_2^0 becomes 2H_2^+ , which involves the loss of 4 electrons ($-4e$). The hydrogen is oxidized. On the other hand, O_2^0 becomes 2O^{--} , which involves the gain of 4 electrons ($+4e$). The oxygen is reduced. We come to the conclusion that *oxidation involves a loss of electrons and reduction a gain of electrons*. Or, we may say that oxidation involves an increase in the positive valence or a decrease in the negative valence ; whereas reduction involves a decrease in the positive valence or an increase in the negative valence.

Notice in this reaction several other facts. When the hydrogen loses four electrons, the oxygen gains four electrons. In every equation, the electron lost by one element is gained by the other. There is no absolute loss, the net result being always zero. Furthermore, a reducing agent, such as hydrogen, loses electrons in the course of a reaction, the hydrogen becoming oxidized ; an oxidizing agent, such as oxygen, gains electrons in the course of a reaction, the oxygen becoming reduced.

242. Copper oxide and hydrogen. Consider another common reaction studied under hydrogen :



Rewriting this equation to show the valences :



We say that the copper oxide is reduced to copper and the hydrogen is oxidized to water. We call copper oxide an oxidizing agent and hydrogen a reducing agent. Now notice that in the change which CuO undergoes the Cu^{++} becomes Cu^0 , but the O^{--} on the left side of the equation still remains the O^{--} on the right side ; it has undergone no change in valence. The Cu^{++} in changing to Cu^0 has gained two electrons. It is therefore an oxidizing agent. The

CuO has therefore been reduced. The H_2^0 has changed to H_2^+ , involving the *loss* of two electrons; which means that the hydrogen has been oxidized.

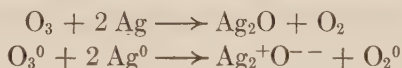
SOME OXIDIZING AGENTS

243. Ozone, O_3 , is, in reality, a form of oxygen which is very active, more active than oxygen itself. Silver, for example, is unaffected by oxygen, even for a considerable length of time. Ozone, however, when brought in contact with silver converts the latter into brown silver oxide.

Ozone may be prepared by passing air or oxygen between two plates charged at high voltage ("silent electrical discharge"). In all oxidizing reactions the ozone is broken down thus:



and it is the O or *atomic* oxygen which gives ozone its powerful oxidizing properties. To take the above example,

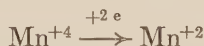


Part of the O_3 in changing to Ag_2O has gained electrons.

244. Manganese dioxide, MnO_2 . In the reaction



the Mn in MnO_2 has a valence of + 4, whereas the Mn in MnCl_2 has a valence of + 2; or



In other words, the manganese *gains* electrons, and the MnO_2 is an oxidizing agent.

245. Nitric acid, HNO_3 . In the equation



the essential part of the reaction is the change of nitric acid to nitric oxide:





In other words, the nitrogen *gains* electrons, and the HNO_3 is an oxidizing agent.

246. Concentrated sulphuric acid, H_2SO_4 . In the equation



the essential change is the conversion of sulphuric acid to sulphur dioxide.

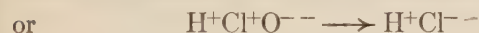


In other words, the sulphur *gains* electrons, and the H_2SO_4 is an oxidizing agent.

247. Chlorine. When chlorine dissolves in water, it becomes an oxidizing agent :



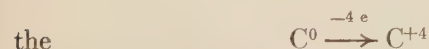
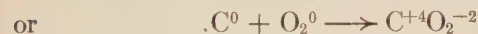
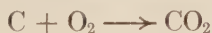
The oxidizing agent is the HClO , or hypochlorous acid. In the course of oxidation the HClO is changed to $\text{HCl} + \text{O}$



The chlorine gains electrons.

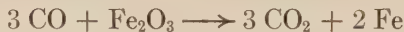
SOME REDUCING AGENTS

248. Carbon, C. In the reaction of carbon and oxygen :

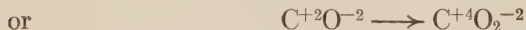


The carbon then *loses* electrons, which is what happens to a reducing agent.

249. Carbon monoxide, CO. In the reaction



the CO is changed to CO_2 . The carbon in CO has a valence of 2, whereas in CO_2 it has a valence of 4.



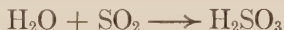
The carbon loses electrons.

250. Hydrogen sulphide, H_2S . Consider the reaction

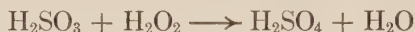


Here the H_2S is changed to S. There is no change in so far as the hydrogen is concerned, but the sulphur loses electrons ($\text{H}_2^{+}\text{S}^{-} \longrightarrow \text{S}^0$).

251. Sulphur dioxide, SO_2 . When sulphur dioxide is dissolved in water, it forms sulphurous acid:



It is the sulphurous acid which acts as the reducing agent. For example:



The sulphur loses electrons.

SUPPLEMENTARY READING

S. R. Brinkley, "Application of the Electron Concept to Oxidation-Reduction Reactions in General Chemistry," *Journal of Chemical Education*, vol. II (1925), p. 576.

E. R. Jette, *Oxidation-Reduction Reactions in Inorganic Chemistry* (The Century Co., 1927).

QUESTIONS

1. Define oxidation and reduction in terms of the electron conception.
2. Select three oxidizing agents and show by means of chemical equations how they react, using the electron concept.

3. Select three reducing agents and show by means of chemical equations how they react, using the electron concept.

4. Make a diagram of the magnesium and oxygen atoms and show, in terms of the electron theory, what happens when magnesium is oxidized.

5. Write an equation showing the transfer of electrons when ferric oxide unites with carbon.

OPTIONAL QUESTIONS

6. In the reaction $\text{NH}_3 + 2 \text{O}_2 \longrightarrow \text{HNO}_3 + \text{H}_2\text{O}$, show, using the electron conception, that the ammonia is a reducing agent.

7. In the reaction $2 \text{N}_2\text{O} + \text{S} \longrightarrow \text{SO}_2 + 2 \text{N}_2$, show that the N_2O is an oxidizing agent. Which is the reducing agent and why?

CHAPTER XX

CHEMISTRY IN AGRICULTURE

252. Out of the air and out of the soil the plant derives its nourishment. It draws carbon dioxide and moisture from the atmosphere and various mineral salts and water from the soil; and with the help of the green-colored pigment, chlorophyll, and certain rays from the sun, the fats, carbohydrates, and proteins are synthesized (built up). These, in turn, serve as the basis for the food of mankind. The carbon dioxide, the moisture, and the mineral salts supply the plant with certain indispensable elements without which no synthesis can take place. A lack of one or more of these elements may starve and destroy the plant just as surely as a lack of certain elements in man's diet will destroy him.

253. Essential elements. Ample experience and many experiments have shown that the following ten elements are essential to the growth of the plant: carbon, hydrogen, oxygen, phosphorus, potassium, nitrogen, sulphur, calcium, iron, and magnesium. Some of these, as has just been stated, the plants get from the air surrounding them; but the majority come from the soil. It is quite probable that a number of other elements, not included in the ten given, will be found, sooner or later, to come under the heading of "essential elements"; but the actual amounts of them required seem to be so small that they have so far escaped attention. Experiments by Bertrand in France, for example, seem to point to the profound influence upon the growth of plants that minute quantities of the element manganese seem to exert.

One thing it is important to bear in mind: the plant, like the animal, cannot make use of elements as such; they must be in the

form of simple, soluble compounds. The exception to this rule is oxygen, which, in man and animals alike, can be absorbed as oxygen for purposes of respiration. But even here, for the purposes of building up plant materials, whatever oxygen is needed comes not from the free oxygen of the air, but from the oxygen which is combined either with hydrogen (in water) or with carbon (in carbon dioxide). The plant foods, most of which are absorbed from the soil by the roots, are, in the main, inorganic in nature.



FIG. 190. The field in front of the man has not been fertilized.

Carbon dioxide and moisture are absorbed by the plant through the leaves, but the larger portion of food is absorbed through the roots from the soil.

254. The soil. The portion of the earth's crust which can support vegetation is called the "soil." In composition it consists of fine and coarse particles of rock in all stages of decomposition, which contain many of the elements necessary for plant growth; and, in addition, organic matter from decayed plants and animals, water, bacteria, fungi, etc. The organic matter is the chief source of the nitrogen.

255. Photosynthesis. While we know that for the growth and development of the plant, light, chlorophyll, and proper temperature are necessary, very much of the process of building up plant structures and plant foods still remains a mystery. We know, for example, that carbon dioxide gas enters the leaf from the outside air by way of the stomata. We believe that this carbon dioxide meets the water carried up from the soil, and, in the presence of light and in the presence of the green pigment, chlorophyll, produces sugar and releases oxygen, thus :



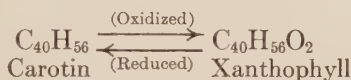
This process in the plant which has just been described is called *photosynthesis*.

Out of the relatively simple sugar, glucose, we suppose, other carbohydrates like starch and cellulose are built; and out of carbohydrates the fats and waxes are built; and there is a theory that from simple sugars and simple nitrogen compounds (obtained from the soil) the proteins are formed.

256. Chlorophyll. The important part played by the green plant pigment in photosynthesis makes it necessary that a word be said about this substance, which is called *chlorophyll*. Despite the work of many chemists over many years, it is only within quite recent times that the chemist has succeeded in isolating pure chlorophyll and analyzing it. It has been found that there are really two varieties of this chlorophyll, one called chlorophyll A, and having the complicated composition $\text{C}_{55}\text{H}_{72}\text{O}_5\text{N}_4\text{Mg}$, and the other chlorophyll B, with the equally complicated composition $\text{C}_{55}\text{H}_{70}\text{O}_6\text{N}_4\text{Mg}$. You will notice that both varieties contain the element magnesium. This is most interesting, for it had long been supposed that chlorophyll, like the hemoglobin of the blood (with which it was thought to be related), contained the element iron. Pure chlorophyll, however, does not contain a trace of iron.

That chlorophyll plays a very important rôle in the respiration of plants, is well known; but it has recently been shown that

together with this pigment are associated two other pigments, also important in respiration; one is carotin, $C_{40}H_{56}$, a substance identical with the pigment formed in carrots and probably the yolks of eggs; and the other is xanthophyll, $C_{40}H_{56}O_2$, a yellow pigment. You will notice that the formulas of xanthophyll and carotin differ by O_2 . It is perhaps worthy of mention that these two substances are very easily transformed into one another, the carotin being easily oxidized to xanthophyll and the latter being easily reduced to carotin:



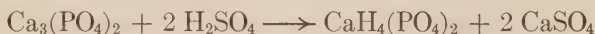
An attractive theory, now current, has it that the oxygen of the air first attaches itself to the carotin, thereby becoming xanthophyll, and that the latter next passes its oxygen on to chlorophyll, the xanthophyll thereby being reduced to carotin.

257. Elements in which soil is deficient. Experience has shown that among the ten elements absolutely essential to the growth of plants three are often found in quantities insufficient for normal development — phosphorus, potassium, and nitrogen. These elements, in the form of their compounds, have to be supplied to the soil if abundant and vigorous crops are desired. The need for these elements has given rise to one of the big industries of the world — the “artificial fertilizer industry.” Let us consider each of these three elements in turn.

258. Phosphorus. Liebig, as far back as 1840, proposed treating bones and rock phosphate with sulphuric acid to make the phosphorus more soluble. This process has been used with but slight modification since his day, and it accounts for the use of about 50 per cent of the total sulphuric acid consumed in the industries.

Basic phosphate slag, obtained as a by-product in the manufacture of steel, has also served as a valuable source of phosphorus.

The chief source of phosphorus for artificial fertilizer is phosphate rock, found in enormous quantities in Florida, Alabama, South Carolina, etc. Equally important deposits of natural phosphates are found in North Africa. The phosphate rock is essentially calcium phosphate, $\text{Ca}_3(\text{PO}_4)_2$, which, being practically insoluble in water, would be but slowly absorbed by the roots if applied to the soil. To make it more readily available, it is usually treated with a limited amount of sulphuric acid, whereby calcium acid phosphate is formed :



Both the calcium acid phosphate and the calcium sulphate are used as fertilizers, the mixture going under the name of "superphosphate."

Incidentally, in this process, sulphur is also supplied to the soil. Professor Lipman, at the New Jersey Agricultural Station, is conducting investigations which seem to indicate that soil micro-organisms have the power of converting ground sulphur in the soil to sulphuric acid. If, then, sulphur and calcium phosphate are mixed and applied to the soil, a certain amount of superphosphate of lime is formed.

259. Potassium. Much of the potassium needed in artificial fertilizers comes from Stassfurt, Germany, and parts of Alsace, France. In these places the potassium is found in the form of a double salt, known as carnallite, with the composition



and from this carnallite, potassium chloride, KCl , may be obtained, which can be used as such in the fertilizer industry.

During the late war, when potash supplies from Germany became unavailable, American chemists made strenuous efforts to locate suitable deposits in this country. While potassium is by no means an uncommon element, it is usually found in such combinations as to make its extraction a costly and laborious process. However, successful methods have now been developed for getting

potash from such sources as the salt lakes in California, the fumes from cement kilns, plant ashes, sugar beet molasses, and the kelp beds on the Pacific Coast.

260. Nitrogen. At first sight it might seem strange that there should be any nitrogen problem at all, seeing that approximately 80 per cent of the atmosphere covering the earth's surface is nitrogen gas. Unfortunately for us, nitrogen as such can no more be utilized by the plant than can nitrogen by the animal. We also need nitrogen. A very characteristic element of all protoplasm is nitrogen, but we can make use of the element in combined form only, — in the form of one of its compounds, the proteins, or their decomposition products, the amino acids. In the case of plants, the nitrogenous compounds which they can turn to account are, chemically, of much simpler type than those necessary for man. They may be either in the form of ammonium compounds, or in the form of nitrates, or in the form of animal refuse, which includes nitrogenous compounds undergoing putrefaction. Let it be emphasized again that, with one exception, all elements necessary for plant and animal life must be presented in the combined, or compound, form before absorption and utilization is at all possible.

261. Chile nitrate. In the early days when fertilizers first came into use, the deficient nitrogen was largely supplied by organic refuse, such as manure. Then later the discovery of nitrate beds in Chile gave rise to an industry not unlike the oil industry of to-day. Chile saltpeter, or sodium nitrate, NaNO_3 , still supplies much of the world's needs for nitrogen fertilizer; so does the ammonia obtained as a by-product in the distillation of coal; but these resources are being tapped so generously that we shall soon reach the end.

262. Nitrogen fixation. The picture of possible starvation and death due to lack of available nitrogen was first graphically painted by Sir William Crookes, the noted chemist, more than a generation ago. He urged chemists to turn their attention to the subject of "nitrogen fixation" — the method by which the nitrogen of the

air can be "fixed" or combined with one or more elements so as to make it available for plant life.

263. The arc process. What are the principles underlying the methods by which the nitrogen of the air can be fixed and so utilized by the plant? The oldest method, the "arc" process, is



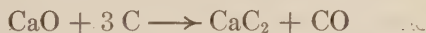
Underwood and Underwood.

FIG. 191. Nitrogen fixation in a lightning storm.

really based on a process which occurs in nature whenever there are electrical storms or lightning outbursts. The nitrogen, present in such abundance in the air, is an extremely inert gas; under ordinary conditions it seems to show little tendency to combine with other elements. Under the influence, however, of an electrical discharge, some of the nitrogen combines with the oxygen (which is also present in the air) and with hydrogen (a con-

stituent of moisture) to form nitric acid. In the arc process, in the place of Nature's electricity, the chemist generates electricity in the form of powerful electric arcs, with much the same results. The nitrogen peroxide which is formed is dissolved in water to give nitric acid. The nitric acid which is obtained is neutralized with lime and the resulting compound (calcium nitrate) can be spread upon the soil just as well as the Chile saltpeter (sodium nitrate) which is mined in Chile. While the raw materials (air and water) needed are cheap enough, the power requirements are not; and though praiseworthy pioneer attempts were made by Bradley and Lovejoy to develop the process at Niagara Falls as early as 1902, it has been in Scandinavia, where water power is, comparatively speaking, very cheap, that the process has met with success.

264. The cyanamide process. The second method is of more than passing interest to us in America, since the Muscle Shoals plant was equipped to use this process. The method is commonly known as the "cyanamide" process. Quicklime and coke are caused to combine in the electric furnace to form calcium carbide, — a substance which is also extensively used for the production of acetylene.



This calcium carbide is next heated in an atmosphere of pure nitrogen (obtained from liquid air); the carbide and the nitrogen combine to form calcium cyanamide,



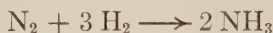
which may be directly spread on the soil, or first treated with water to liberate ammonia,



and the latter combined with sulphuric acid to form ammonium sulphate, — an excellent fertilizer. Air, water, limestone, coke, and coal are the raw materials. The electrical power requirements are approximately one fourth those required for the arc process; so

that, for a country such as our own, the cyanamide method deserves attention. There are, however, drawbacks. Despite its comparative cheapness, the cyanamide is an unpleasant substance to handle and to work with, and is somewhat poisonous, both to the plants and animals.

265. The Haber process. By far the most successfully developed method for the fixation of nitrogen is the Haber, or "synthetic ammonia" process. It has been developed to such a degree in Germany, that about one half of the total world's supply of "fixed" nitrogen is obtained by its means. In theory the process is simplicity itself; nitrogen and hydrogen are caused to combine with one another, forming ammonia:



In practice, the chemical, and particularly the engineering, difficulties are many and difficult to overcome. The pressure and temperature have to be very carefully regulated; and for the process, a catalyst is needed.

Haber and his associates, after many trials, found a suitable catalyst. Its exact chemical composition is not known, though there is reason for believing that it is probably some iron compound containing traces of other elements. As the key to the process is the discovery of a suitable catalyst, it is most gratifying to learn that the scientists of our own Fixed Nitrogen Research Laboratory have developed their own catalyzer, which is said to be even more efficient than the one employed by the Germans.

A very promising development in recent years has been the attempt to manufacture a compound containing both nitrogen and phosphorus in a form available to the plant, namely, ammonium phosphate, $(\text{NH}_4)_3\text{PO}_4$. This is done by first making phosphoric acid and then combining the latter with ammonia (obtained by the Haber process). The phosphoric acid is prepared by strongly heating rock phosphate with sand and coke, using the electric furnace for this purpose.

At the beginning of the twentieth century, the principal producer of nitrogen compounds was Chile. The only other source was the ammonia obtained from the destructive distillation of coal. In 1905 synthetic nitrogen compounds began to be made in Norway by the arc process and in Germany and Italy by the cyanamide process. In 1913 the Haber process was introduced in Germany.

In 1913, 56 per cent of the world production of nitrogen was provided by Chile saltpeter, 37 per cent by nitrogen compounds obtained from coal, and only 7 per cent by synthetic nitrogen compounds. In 1925 only 30 per cent of the total production came from Chile saltpeter and 44 per cent from synthetic nitrogen compounds! Can more eloquent tribute than this be paid to the achievements of the chemist working in his laboratory?

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F. E. Allison, "Nitrogen as a Plant Food," *Journal of Chemical Education*, vol. III (1926), p. 50.

J. G. Lipman, "The Value of Sulphur in Soil Improvement and Rock Production," *Industrial and Engineering Chemistry*, vol. XVI (1924), p. 250.

J. G. Lipman, "The Nitrogen Problem in Agriculture," *Journal of Chemical Education*, vol. IV (1927), p. 845.

QUESTIONS

1. What are some of the essential elements necessary for plant growth, and in what form must these elements be before they are made use of?

2. Explain the word "photosynthesis."

3. What is chlorophyll and why is it important in plant growth?

4. In what form is phosphorus applied to the soil?

5. What are some of the sources of potassium?
6. Why is it necessary to supply nitrogen to the soil when there is such an abundance of the element in the air?
7. What do we mean by "nitrogen fixation"?
8. Give brief accounts of the manufacture of nitrogen compounds by three different methods.
9. Why did not the early settlers of this country pay any attention to fertilizers?
10. How many tons of sulphuric acid 65 per cent pure are needed to produce 100 tons of calcium acid phosphate?
11. How many liters of hydrogen are needed to combine with 50 liters of nitrogen to form ammonia? How many liters of ammonia are obtained?
12. Wood contains the elements carbon, hydrogen, and oxygen. What products are obtained when the wood is burned?

OPTIONAL QUESTION

13. What products would you expect from the decomposition of wood, after it had been covered completely by mud?

CHAPTER XXI

ROCKS, GLASS, AND CERAMICS

266. As we have already seen, in the chapter on "Earth," this planet we live on is made of rocks, covered over with a thin layer of loose material called earth. Most rocks contain the element silicon, which, next to oxygen, is the most abundant in nature.



FIG. 192. Earth and rock strata.

Roughly speaking, there is about 50 per cent oxygen and 25 per cent silicon. Granite, the feldspars, mica, hornblende, slate, sandstone, and many others are common rocks and minerals which are silicates of sodium, potassium, calcium, iron, aluminum, and magnesium. In fact, the only common mineral which is not

a silicate is limestone, CaCO_3 . One of our most common minerals is silica, SiO_2 , which occurs in many forms as shown in the following table.

FORM OF SiO_2	DESCRIPTION
Quartz (crystal)	transparent white crystalline
Amethyst	transparent bluish violet crystalline
Carnelian	transparent red crystalline
False topaz	transparent yellow crystalline
Opal	translucent
Chalcedony	translucent waxy
Chrysoprase	green chalcedony
Sard	brownish red chalcedony
Agate	banded in different colors
Onyx	like agate — straight bands
Flint	opaque — gray to brown
Jasper	opaque red, yellow, etc.; smooth surface
Jade	green; smooth
Sand	
Sandstone	
Infusorial earth	fine white powder

Silica has about the same hardness as glass and the hardest steel, and for this reason it is used as an abrasive. The sand blast, sand paper, sandstone (used for whetstones), and electro-silicon (infusorial earth) are examples of silica used as abrasives. Flint (SiO_2) and steel were used at one time to make fire. The hard flint knocks off bits of the steel, which, heated by the great friction, take fire.

Many of the forms of silica are rather pretty, especially when cut and polished; they are therefore used in jewelry as semi-precious stones. Vitreosil is fused quartz made into vessels for chemical uses. It has a melting point of 1750°C . and is absolutely acid-proof. Furthermore, it will not crack under changes of temperature, as glass does: a red hot silica dish may be plunged

into water without cracking. The best camera lenses are made of quartz, which transmits ultra-violet light much better than glass; and for this reason several kinds of lamps used for medical purposes are made of quartz.

267. Weathering of rocks — formation of soil. As before mentioned most rocks are silicates of a few elements: Na, K, Ca, Mg, Al, Fe. Now what happens to such material under the action of weathering — that is, a slow, but constant attack of oxygen, water, and carbonic acid? Sodium and potassium form carbonates, while calcium and magnesium form bicarbonates, all of which, soluble in water, are ultimately carried into the ocean. Iron forms ferrous bicarbonate, soluble in water, but on oxidation it is precipitated as ferric hydroxide $\text{Fe}(\text{OH})_3$, a red-brown powder, which colors nearly all rocks and soils. Aluminum silicate, unaffected by natural weathering agents, remains as clay. Silica is also insoluble and remains as such. Sometimes water carries off the clay in suspension, leaving the sand (silica); while in other places, when the water becomes quiet, the clay settles, and may form slates and shales. Sand and clay, then, are the remains of the chemical weathering of rocks, and this mixture is the chief constituent of earth or soil. Sand gives the soil porosity and clay gives it plasticity and other properties necessary for the growth of plants. The chief use of silica and the silicates is in the manufacture of glass, glazes, pottery, earthenware, and enamels.

268. Glass. When silica is fused with soda (Na_2CO_3), it forms a clear transparent mass, sodium silicate, which, on cooling, resembles glass:



This sodium silicate is soluble in water and is therefore called *water glass*, which is used to preserve eggs. (It closes the pores and so prevents microorganisms from entering.) It is also used to make cements, for fireproofing wood and cloth, and as a filler in soaps. Both sodium and potassium silicates are colloidal, and therefore do not crystallize on cooling.

Glass is a gelled mixture of an alkali silicate (Na or K) with one or more other silicates, and silica. The alkali silicate acts as a protective colloid, preventing the crystallization of the calcium or lead silicate which would make the glass less uniform. Very old panes

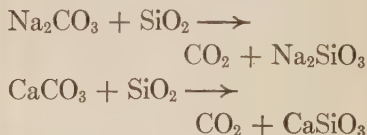


Ewing Galloway.

FIG. 193. Glass blowing.

of glass sometimes do crystallize and lose their transparency. Now, while glass contains two or more silicates, it is not made by simply mixing these silicates; but (if window glass be taken as an example) by fusing together pure sand, SiO_2 , soda, Na_2CO_3 , and limestone, CaCO_3 . The formation of the silicates may be considered as taking place according to the following equations:

of glass sometimes do crystallize and lose their transparency. Now, while glass contains two or more silicates, it is not made by simply mixing these silicates; but (if window glass be taken as an example) by fusing together pure sand, SiO_2 , soda, Na_2CO_3 , and limestone, CaCO_3 . The formation of the silicates may be considered as taking place according to the following equations:



the evolution of the carbon dioxide tending to stir the mass. When the molten glass has cooled to the proper consistency, it is shaped by blowing, pressing, or rolling. Plate glass is made by pouring the

gel on an iron table and rolling to the proper thickness. The surface, quite rough after rolling, is polished by rubbing with finer and finer grades of sand, and finishing with rouge, an exceedingly fine powder. In making tumblers, lampshades, and similar glass articles, the molten glass is shaped in a mold by means of a

plunger, shaped like the inside of the article. Bottles, vases, etc. are made by blowing the glass with a long trumpet-shaped pipe, sometimes in a mold, managed by another workman. Window panes are made by blowing long bubbles, rotating on a table, and then cutting off the ends and slitting the cylinder. On heating, this flattens out. Simple shapes, like window panes, bottles, and test tubes, are now made by glass-blowing machines.

269. Annealing. Glass being a poor conductor of heat, rapid and irregular cooling will cause stresses, which make the finished article exceedingly fragile. This may be remedied by annealing, or slow cooling, during which the finished glass articles are moved slowly through a heated atmosphere, sometimes as much as 300 feet.

270. Kinds of glass. The most common varieties of glass are crown, Bohemian, flint, and Pyrex. *Crown* glass is a sodium-calcium mixture used for window and plate glass, bottles, and other ordinary articles. It is easily fused, soft, and slightly soluble in water. *Bohemian* glass is a calcium-potassium mixture, and is harder, less fusible, and less soluble than crown glass. It is used for vases, tumblers, and the better grade of chemical glassware. *Flint* glass is a potassium-lead mixture, heavy, soft, and brilliant. It is used for "cut-glass," imitation jewelry ("paste"), and optical lenses. *Jena, non-sol* and *Pyrex* glasses are all made with oxide of boron as well as silica, containing, therefore, a mixture of borates and silicates. They are hard and infusible, quite insoluble, and stand temperature changes better than ordinary glass.

271. Armored glass contains nickel-steel wire, which, in case of breakage, will support the pieces. Nickel-steel has the same low heat expansion as glass.

Glass for automobile windshields is made up of two plates of glass with a sheet of celluloid in between. This same glass becomes bullet-proof when the thickness is increased.

272. Color in glass. The green color of ordinary cheap glass is due to the presence of ferrous compounds, formed by the rust in

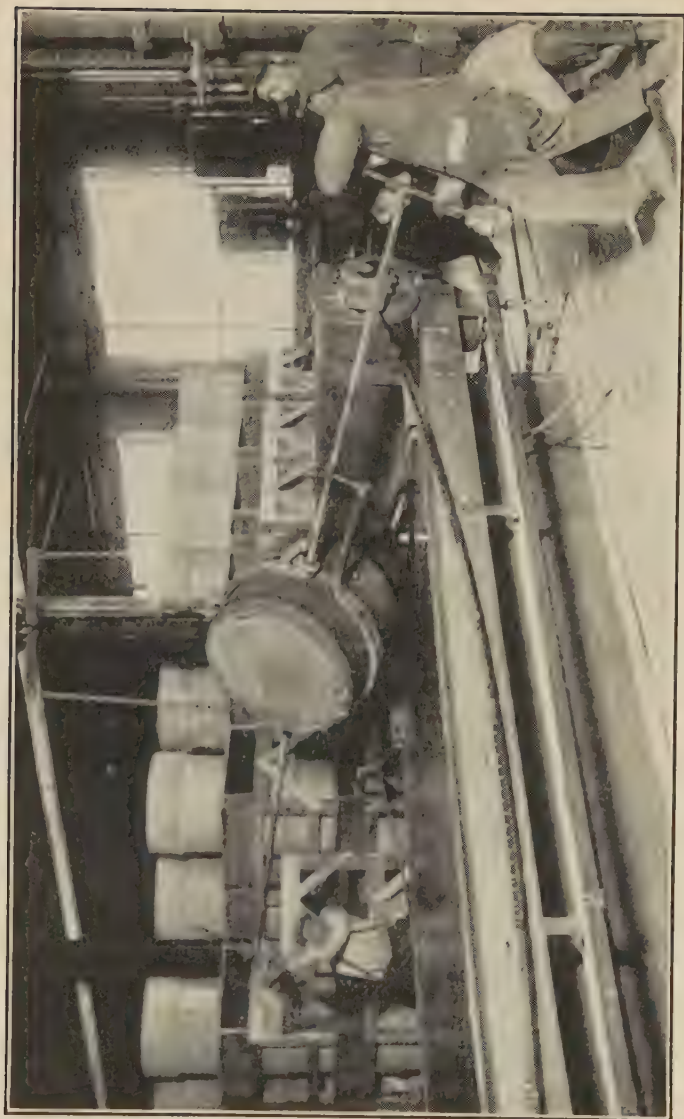


Fig. 194. Making plate glass.

the sand or limestone. It may be decolorized by manganese dioxide, which oxidizes it to ferric compounds whose color, a light yellow, is negligible. White glass owes its color to a suspension of some rock, like feldspar, or cryolite. As we have seen in the chapter on colloids, colors are at a maximum in the colloidal condition; and since we are dealing with sodium silicate, a protective colloid, there is a tendency to form colored substances in fused glass. Ruby glass owes its color to colloidal gold, copper, or selenium; amber glass owes its color to silver, sulphur, or carbon; green glass to chromium or copper; blue glass to cobalt.

273. Stained glass is ordinary glass with a thin layer of colored glass on the surface.

274. Enamels are essentially glasses of special composition.

In the case of enameled cooking utensils, nickel-steel is usually the base, while the glasses must be free of lead and zinc, whose compounds are poisonous. The article to be enameled is dipped into a paste of ground glass and water, and is then baked until the glass melts. Enameled bath tubs and sinks are cast iron coated with a white lead glass.

275. Ceramics. This name is given to products of the potter's art: objects molded from clay. Clay is a colloid and becomes plastic during gelation (coagulation). This state may be maintained indefinitely; but on heating, an irreversible change takes place and crystallization sets in, causing the hardening of the mass. Bricks, tile, sewer pipe, terra cotta, earthenware, and porcelain are all ceramics. Bricks are made of impure clays, the red color being caused by the high iron content. Tile, porcelain, and china are made of pure white clays (kaolin).



FIG. 195. Enameled ware.

276. Glazes. These are essentially glasses. The glaze on stoneware is produced by throwing salt (NaCl) into the fire, forming the glass sodium silicate, plus other silicates, on the surface. The surface of porcelain resembles crown glass, and is made of lime, clay, feldspar, and quartz.



FIG. 196. Porcelain ware.

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 H. E. Howe and F. M. Turner, *Chemistry and the Home* (Charles Scribner's Sons, 1927), p. 67.

QUESTIONS

1. Mention some common minerals and tell what elements each contains.
2. How does amethyst differ from quartz?
3. How could you recognize opal?
4. Why is quartz glass expensive?
5. Explain how "earth" is formed. Why is it chiefly sand and clay?
6. What is glass?
7. How does age affect glass?
8. How is glass annealed? In what way is such glass superior?
9. Make a table giving in a vertical column the kinds of glass, and in a parallel column, one article (not mentioned in the text) made of each kind of glass.
10. What is the difference between the enamel on a bathtub and that on a kitchen utensil?
11. Name a number of articles, not mentioned in the text, which are made of clay.
12. How does a glass differ from an enamel?
13. Make a list of articles which are essentially silica, and with which you, personally, are familiar.
14. Why does some glass crack easily?

15. How is blue glass made? What is ruby glass used in "dark rooms"?
16. How is the colored glass made that is sometimes seen in churches?
17. How many kilograms of pure sand would be needed to make 100 kg. of water glass?
18. How is glass etched? Write the equation.
19. Write the formula of the silicon compound corresponding to CO_2 ; to Na_2CO_3 .
20. In the equation for glass making, what volume of carbon dioxide is evolved when 1 kilogram of sodium carbonate is used?
21. What is the percentage composition of silicon dioxide?

OPTIONAL QUESTIONS

22. Recalling your studies in general science, show how the roots of plants are adapted to growth in soils containing clay and sand.
23. How would you separate each substance from a mixture of sugar, ground chalk, and sand?
24. Why do plants grow better in soils containing both sand and clay?
25. Referring back to the chapter on colloids, suggest tests to be employed to show that sodium silicate is in the colloidal condition.
26. How would you expect sodium silicate to react toward litmus? Why?

CHAPTER XXII

FUELS AND ILLUMINATION

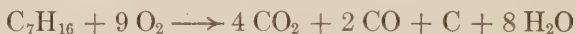
277. The heat and light supplied to us by the sun, and without which life would be impossible, is supplemented by us from other sources, from coal and petroleum, for example; though it must not be forgotten that coal and petroleum owe their fuel value to energy originally obtained from the sun.

The fuels commonly used may be divided into (1) solid fuels, such as coal, coke, wood, charcoal; (2) liquid fuels, such as kerosene, gasoline, crude oil — all three obtained from petroleum — and alcohol; and (3) gaseous fuels, such as coal gas, natural gas and acetylene. A characteristic of all the fuels used is that, when burned, giving off heat, and, in some cases, light, it is the carbon and the hydrogen of the fuels which are involved in the chemical process, the carbon being oxidized to carbon dioxide and the hydrogen to water. Incomplete combustion may give rise to the harmful carbon monoxide, or some of the carbon may escape as such, forming soot or smoke. This deposit of carbon, for example, is the source of great annoyance to motorists, inasmuch as the carbon deposits in the cylinders.

The completeness of the combustion of the fuel varies with the amount of oxygen and the uniformity of the mixture. To illustrate what may happen, let us assume that gasoline has the formula C_7H_{16} ; then, in complete combustion, we have:



in incomplete combustion we may get:



278. Coal. There are two varieties of coal, anthracite or hard coal, and bituminous or soft coal. Representing as it does fossil-

ized wood, it is no wonder that coal should be found to consist of volatile compounds of carbon and hydrogen (hydrocarbons), mineral matter and moisture, in addition to the very large percentage of carbon. From the point of view of fuel value, the ash and moisture are useless. The essential difference between hard and soft coals is that the former contains a large percentage of free carbon and a relatively small percentage of hydrocarbons, whereas the latter contains relatively little free carbon and much hydrocarbons. Being cleaner to handle, and producing less smoke, the hard coal is preferred for household purposes. On the other hand, since soft coal usually contains less ash than hard coal, more heat is obtained from the former. Heat, as you may know, is measured in calories. Large users of coal buy the fuel on the basis of the number of heat units per ton; the householder, unfortunately for him, cannot at present do so, and thereby often gets a coal which is found to be of inferior quality only after he has used it in his fireplace or furnace.

279. Coke is the residue left when coal is heated in retorts in absence of air — *destructive distillation*. It contains enough heat value to be used as a fuel, though its main use, in this country at least, is as a reducing agent in the blast furnace.

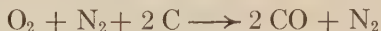
280. Wood is used as a fuel only in certain parts of the country where timber is cheap and abundant; but, of course, it is commonly used to start fires.

281. Charcoal, or wood charcoal, made by heating wood in closed retorts, is sometimes used as a kindling material, or for fireplaces.

282. Gasoline, a product of petroleum, is the chief fuel used for automobiles and other motors. *Kerosene*, another petroleum product, is still used very largely as an illuminant (in lamps) and as a fuel in farm tractors. *Alcohol*, in addition to its many other uses, is sometimes used in the place of gasoline, and will probably be used more and more as the gasoline becomes scarcer.

283. Producer gas. A limited supply of air passed through incandescent coke or coal yields a mixture of carbon monoxide and

nitrogen known as *producer gas*. If air, for the sake of convenience, be written $O_2 + N_2$, then we can write :



This producer gas is used in the industries as a fuel. Since coal with a high percentage of ash can be used in its manufacture, — a percentage of ash which would make the ordinary burning of such a coal very difficult, if not impossible, — producer gas is coming to be more and more widely used.

284. Water gas is made by passing steam through incandescent coke or anthracite coal :



the mixture of carbon monoxide and hydrogen being spoken of as *water gas*. Both the gases burn and a large amount of heat is,

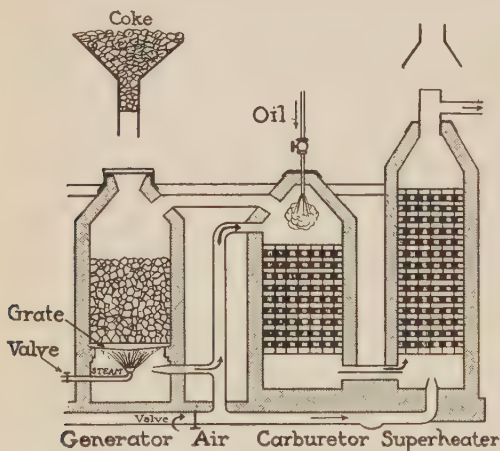


FIG. 197. Manufacture of water gas.

therefore, evolved. It finds extensive use as a fuel and in the manufacture of illuminating gas. In fact, more than 50 per cent of the illuminating gas used in this country is water gas. It is now the chief source of industrial hydrogen. This is obtained by passing water gas and superheated steam over iron oxide (a

catalyst), whereby the carbon monoxide and steam combine :



the carbon dioxide being removed by lime.

By *enriching* the water gas — that is, by mixing it with oil vapor — the gas can also be used for illuminating purposes.

285. Coal gas, an important constituent of the common *illuminating gas* of our household, is manufactured by the destructive distillation of coal. The coal is strongly heated in retorts made of firebrick, and the vapors evolved first pass through pipes in which the tar condenses; then through “scrubbers” containing water

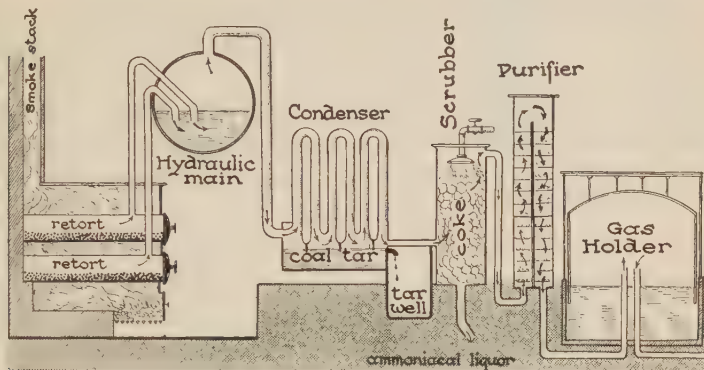


FIG. 198. Manufacture of coal gas.

which absorbs the ammonia; then through a “purifier” containing iron oxide which removes hydrogen sulphide, etc.; and finally into the “holder” in which the gas is stored. What remains in the retort is coke.

You will notice that the principal products obtained from the destructive distillation of coal are coke, coal tar, ammonia, and coal gas, all highly valuable substances. One of these products, coal tar, was at one time considered a mere nuisance and was thrown away; now most of the dyes and drugs on the market are made from it.

Coal gas consists, among other things, of about 30 per cent methane, CH_4 , and about 50 per cent hydrogen.

286. Natural gas, often found in localities rich in petroleum, is obtained direct from the earth by driving small holes deep into the ground. It is rich in methane and other hydrocarbons and is often used for heating purposes (in place of illuminating gas) in those localities near which it is found.

287. Stoves and their regulation. In primitive times man used an open wood fire to warm himself and cook his food, and such remained the sole means of heating, even up to the middle ages.

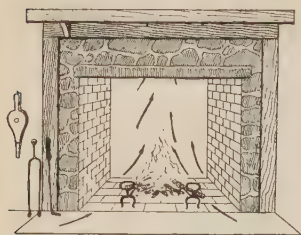


FIG. 199. Open fireplace.

Indoors, the problem of smoke disposal soon led to a hole in the roof, a device resulting in the construction of the open fireplace and chimney. It was soon discovered that the chimney caused a draft, regulation of which by a damper enabled one to exercise much better control over the rate of combustion.

It was found that when the logs were held up off the floor, they burned better; hence andirons. To keep off the excessive heat from those who sat around the fireplace fire screens were devised. Gradually the fire was inclosed more and more and we got a stove.

The modern stove is frequently used both for cooking and heating, and therefore we shall consider only a cooking range. The coal or wood is ignited on a grate which supports it, so that air may enter the burning

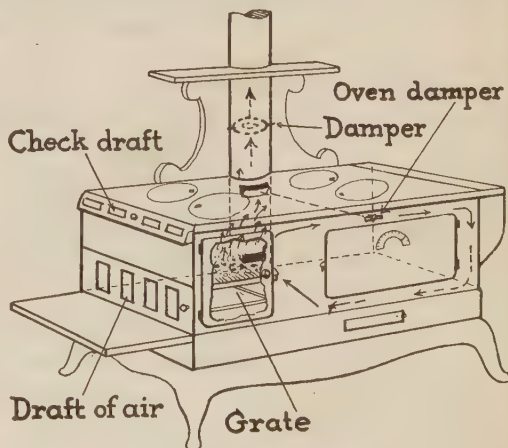


FIG. 200. A kitchen stove.

mass from below, while ashes fall through instead of clogging the fire. The chimney creates a draft, the hot light gases rising, while the air from the outside rushes in to take its place. Both these

currents of air are controlled. The damper on the chimney may be turned so that the opening is smaller, which then hinders the escape of the hot gases.

A slide door on the stove under the grate admits more or less air from the outside, thus controlling the draft. Above the fire box another slide door, the check draft, may admit air. This cuts down the amount of air entering the draft. At the same time air admitted above the fire helps to burn completely any unburned fuel gases which would otherwise escape up the chimney. A smoke damper is used to shut the direct opening into the chimney, forcing the hot gases to find their way around the oven and then into the chimney through another opening under the oven. These are the usual controls on the kitchen stove.

Let us study, now, the regulation of a stove for varying conditions.

KIND OF FIRE	DAM- PER	DRAFT	CHECK	OVEN DAMPER	REMARKS
1. Starting	open	open	closed	open	most air, rapid burning
2. Rapid cooking	open	open	closed	open	most air, rapid burning
3. Slow cooking	partly open	shut	partly open	closed	retains most heat in house
4. Quick oven	open	open	closed	closed	
5. Slow oven	partly open	partly open or even closed	closed	closed	
6. Very slow fire overnight	nearly shut	shut	open wide	closed	least air, but gases may escape into chimney

In the early days dampers were made of a piece of flat metal which could entirely shut off the chimney; in which case, the products of combustion found their way through cracks in the stove into the room, and many cases of death through asphyxiation

resulted. (Incomplete combustion yields carbon monoxide.) Most modern stoves have dampers made in such a way that even when entirely closed, the gases may still escape up the chimney through a hole in the damper.

288. How to start a fire. The common method of starting a fire by piling wood upon paper and coal upon wood and first lighting the paper, is based upon much practical experience. Of the three, the paper reaches its kindling point first; which means that the paper will burn more easily than either the wood or the coal. In doing so, heat is evolved. In this way, the kindling temperature of the wood is reached, which, in turn, starts burning. The

wood, in burning, gives off heat; and after the kindling temperature of the coal is reached, it starts burning.

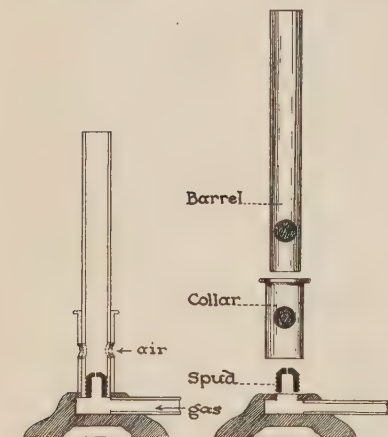


FIG. 201. A bunsen burner.

289. The bunsen burner.

When gases burn, they often produce not only heat but light. Ordinary coal (or illuminating) gas when burned produces a considerable amount of light as well as heat. For purposes of illumination, we want as much light and as little heat as possible. On the other hand, where the gas is to serve as a

fuel, the light emitted by it is of no consequence; what we want is a maximum amount of heat.

Now it can be shown that in many fuels, coal gas included, the luminosity is due to the incomplete combustion of the gas whereby carbon particles are heated until they glow. That there are such carbon particles in a gas which is luminous is shown by putting a porcelain dish into the flame. When we take this dish out of the flame, we find it covered with soot, which is a form of carbon.

Obviously, if we could, in some way, burn up this carbon, we would get additional heat. The luminosity, it is true, would disappear; but for heating purposes, as we have already stated, we want a fuel to give as much heat as possible, and we are not at all concerned with its lighting powers.

The *bunsen burner* accomplishes this object. Here the gas and air are first mixed in a tube before the gas is allowed to burn. The excess air so supplied causes all of the carbon in the fuel to burn, giving therefore more heat than in the ordinary burner. The gas is no longer luminous, and a porcelain dish thrust into the flame will no longer show a black deposit.

The mixing of the gas and air in the barrel is brought about by a stream of gas issuing through the spud, at the same time drawing air through holes in the side.

The gas cooking range employed in kitchens is nothing but a burner in the horizontal position.

290. Illumination. From time immemorial vegetable and animal fats and waxes have been used for illumination. It is only within the last few generations, with the production of kerosene from petroleum, that kerosene lamps have come into use. But the kerosene lamps of the earlier days were often dangerous to handle, due to the unstable and explosive character of the kerosene (which had not been sufficiently refined), and often with an offensive odor — a result due to the presence of sulphur compounds in the oil which, when burned, would evolve evil-smelling sulphur compounds. In time a method was discovered for removing the obnoxious sulphur substances, and a method was also found for producing kerosene which burned quietly.

Next it was discovered that combustible gases with great illuminating power could be obtained and utilized by the destructive distillation of coal. This developed into the enormous gas industry

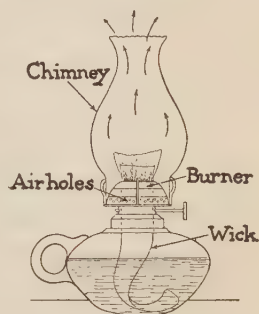


FIG. 202. An oil lamp.

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of the present day. Some notable improvements in more recent times were to mix water-gas with the coal gas, and to enrich the mixture by adding hydrocarbon gases obtained by the "cracking" of oil.

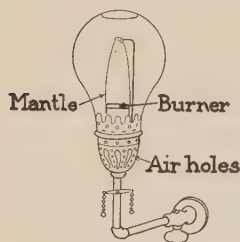


FIG. 203. A Welsbach mantle.

The mantle burner, introduced by Welsbach in 1884, was another step in improving illumination. Here the mantle is impregnated with solutions containing the compounds of two rare elements, thorium and cerium, which possess the property of glowing brightly when heated strongly. The intense heat is obtained by converting the

ordinary gas burner into what is essentially a bunsen burner.

The most recent development, the use of electricity for lighting purposes, does not involve any oxidation process at all (as do the methods so far discussed), but depends upon heating a filament in the absence of oxygen, the heat being obtained by the resistance developed to the passage of electricity, which heat causes the filament to glow and to give light. The filaments first used were made of carbon; these were followed by filaments made of tantalum, and finally by those made of tungsten, the most difficult metal to melt.

The production of tungsten in the form of fine-drawn wire was a great metallurgical and chemical triumph. But equally so was the discovery that instead of using bulbs from which the air had been removed as completely as possible, even better results were obtained by substituting one of the "inert" gases (nitrogen or argon) of the air.

291. Luminosity is caused by the heating of a solid to incandescence. In the ordinary illuminating gas, the luminosity is due

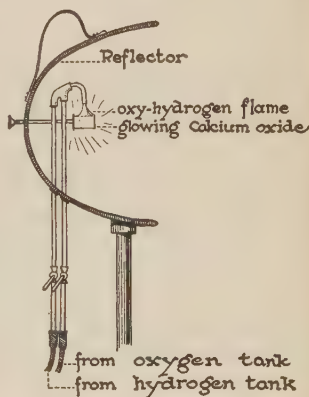


FIG. 204. The calcium light.

to the presence of particles of carbon. In the "calcium light," the burning hydrogen itself produces no light, but when it plays upon the calcium oxide, light is produced. In the Welsbach mantle, it is the heating of the oxides of cerium and thorium to incandescence which gives the light. In the electric lamp the filament is heated to incandescence.

SUPPLEMENTARY READING

- W. H. Fulweiler, "Fifty Years of Gas Chemistry," *Industrial and Engineering Chemistry*, vol. XVIII (1926), p. 945.
 H. E. Howe and F. M. Turner, *Chemistry and the Home* (Charles Scribner's Sons, 1927), p. 274 (illumination).

QUESTIONS

1. Trace the origin of the fuels found in nature.
2. What is a common fuel used (a) in your home, (b) in the automobile, (c) in the lamp, (d) for cutting steel?
3. What products are formed, (a) in the complete combustion of fuels, (b) in the incomplete combustion of fuels?
4. Point out the advantages and the disadvantages of two varieties of coal.
5. Upon what does the fuel value of a coal depend?
6. Name the two liquid and two gaseous fuels commonly met with and give their sources.
7. Describe in your own words what is meant by a fuel.
8. Show how you would regulate the damper, draft, check, and stove damper in a stove (a) when a fire is to be started; (b) when the stove is to be used for rapid cooking; and (c) when the stove is to be used for slow cooking.
9. Account for the dense black smoke so often met with in large cities. Suggest a possible way of eliminating this nuisance.
10. What is the danger of having an automobile motor running in a closed garage?
11. Why is the flame in the bunsen burner non-luminous?
12. What products are formed when the hydrogen and methane in coal gas are burned? Write the equations.
13. Mention three factors necessary to carry on combustion.
14. When the fire in the furnace is almost "dead," the addition of too much coal causes the fire to go out. Explain.

15. Write reactions to show what happens when coal is burned (a) on the grate, (b) in the bed of the fire. What happens above the fire?
16. Mention cases in which hydrogen is used as a fuel.
17. Why cannot fine coal be burned on a grate?
18. One of the constituents of chimney gases is sulphur dioxide. What does this show about the composition of such coal?
19. Soft coal contains nitrogen compounds. What becomes of the nitrogen during destructive distillation?
20. Why is a blue flame produced when fresh coal is added to a fire?

OPTIONAL QUESTIONS

21. Give some uses for the products obtained from the destructive distillation of coal.
22. In view of the fact that wood shavings will burn as readily as paper, why are both paper and wood used in starting a fire?
23. Of what value is the chimney in the running of a furnace?
24. Why does an oil lamp sometimes smoke?
25. How can very fine coal (powder) be utilized as fuel?
26. Starting with 50 liters of gasoline vapor (formula C_7H_{16}), how much air would be required to burn it completely?
27. A certain gas has the following composition: C, 81.82%; H, 18.18%. If 10 liters of this gas weigh 19.80 grams, what is the formula?
28. What is red ash coal?
29. Why are wood ashes thrown on a lawn?

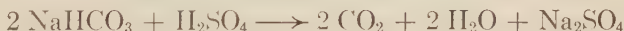
CHAPTER XXIII

EXPLOSIONS

292. We are all familiar with stories of explosions in coal mines, grain elevators, flour mills, and gas plants, and some have even read of explosions in chocolate factories and sardine canneries, but who would imagine the possibility of a disastrous explosion in the process of satin-finishing aluminum? Whenever we light the oven burner of a gas range or strike a match near an automobile, an explosion, fatal in its consequences, may be the result. It is, therefore, important to know: what is an explosion, how is it brought about, and how may it be prevented?

293. Nature of an explosion. When water is changed to steam, one volume of water expands to 1700 volumes of steam. Now in the steam boiler this is done gradually, and the pressure developed is never permitted to get too high (a safety valve letting out the excess steam); but if the water is vaporized suddenly, by coming into contact with a red hot surface, the large volume of steam cannot get out of the safety valve as fast as it is formed, the pressure rises rapidly, and the boiler explodes.

The same thing may take place when a can of soup is placed on a gas flame to heat. The fire extinguisher presents another possible explosion. The action of acid on a carbonate is a fairly rapid chemical change and the pressure developed by the gas produced, carbon dioxide, is made use of to throw a stream of water on a fire. The equation explains this:



Now the large amount of water in the fire extinguisher keeps down the temperature, so that the expansion of the gas is not great

enough to produce an explosion; but if the soda were dry, and confined in a small space, an explosion would result.

Sodium and potassium are very active, and their reactions, particularly with water and acids, are usually explosive. The reaction with water, for instance,



produces hydrogen, which is a gas, and here again, if the amount of water is small, the heat developed raises the temperature sufficiently to cause an explosion, especially if air is present, so that the hydrogen can burn. The equation for the second reaction is $2 \text{ H}_2 + \text{O}_2 \longrightarrow 2 \text{ H}_2\text{O}$, and this is always explosive because of the great heat developed. The reverse equation,



would be even more explosive, because of the great expansion from liquid to gases; but since great heat is produced when the gases combine, it would require the same amount of energy to decompose the product, water, as was produced during its formation. That means, the reaction will proceed only so long as energy is added, and it can therefore never become explosive. An explosive reaction is always one, then, which liberates energy. It may require a little energy to start it, but once started, it proceeds, by itself, to completion. Somewhat like a wagon on the slope of a hill, it may need a little push to get started, but it will run the rest of the way by itself, and in fact, go faster and faster. To sum up, then, *an explosion is a change, involving great heat and rapid expansion of gases.*

In order that combustion may go on, there must be present the combustible and the supporter of combustion and these substances must be kept at or above the kindling temperature. Two substances are necessary, then; these must combine chemically, and in order to do so, they must be in contact. (For example, it is evident that coal in the fire box cannot combine with oxygen out-

side the house; the oxygen must be led into the furnace, and must actually touch the coal.) Now, suppose we have a lump of coal burning. Why doesn't it burn up at once? The inside cannot burn because it is not touching oxygen. Only the outside burns. Suppose the lump were split in two. Then burning could start on a greater surface than before. Splitting the two pieces of coal would add still more surface, and, to sum up, the smaller the pieces of coal, the greater the surface, and the more readily can oxygen combine with it. The large lump of coal burns slowly, then, because the inside portion cannot burn until the outside

Large lumps



FIG. 205.

Small lumps



FIG. 206.

Coal dust and oxygen mixture



FIG. 207.

layer has been burned away. When, therefore, can the coal burn most rapidly? Apparently when it is powdered most finely, and each particle is surrounded by enough oxygen to combine with it. That is one kind of explosive mixture: a thorough mixture of two substances which can combine chemically. Whenever two substances burn, they can also explode. That is, they can burn most rapidly. An explosion of a mixture is a chemical change that liberates energy and which takes place in a very short time; in other words, a very rapid combustion. And why has time such a great effect? Consider the ordinary burning of coal and the explosion of coal dust. The ordinary burning produces new substances, like carbon dioxide and water (as well as heat), which escape as fast as formed. But when these are produced very rapidly, the hot gases cannot escape fast enough and the temperature rises enormously. The increased temperature expands the gases, which, if confined, increase the pressure until

something bursts. Take note, however, that if not confined, the gases will escape, nothing will burst, nor need there even be a sound.

294. Matches. Some of the earlier varieties of the match, like the "lucifer," were tipped with a mixture of sulphur and potassium chlorate, which we recognize as an explosive. When rubbed gently, such a mixture crackles and throws off sparks, and when struck, the entire mass bursts into flame with a loud report. These matches were dangerous, and many disastrous fires were caused by

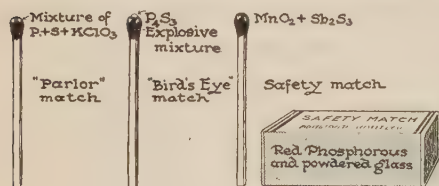


FIG. 208. Matches.

their use. If such a mixture is ignited by heat, without friction, a rapid but quiet combustion takes the place of the noisy explosion. To this end, phosphorus was next tried, which, taking

fire at a lower temperature, set fire to the explosive mixture in a safer way. That did not help much, for it was found that mere rubbing against each other could ignite such matches. Next came the safety match. The head of such a match contains manganese dioxide (a weaker oxidizing agent than potassium chlorate) and antimony sulphide (a less inflammable substance than sulphur, although it is the sulphur in the compound which is the combustible). The match box is painted on the side with a mixture containing powdered glass and red phosphorus (much less inflammable than the white phosphorus formerly used). Such a match can almost never be ignited by accident. The head contains an explosive mixture, it is true, but this is of very low order, requiring a great deal of friction or direct heat to start the explosion. When the match is struck against the mixture on the box, the friction heats up the red phosphorus, which oxidizes but does not burn (the glow can be seen in the dark). The heat developed by the oxidation is sufficient to raise the temperature of the explosive mixture to kindling, and this then bursts into flame, burning very rapidly, but with no report; that is, a slow explosion takes place.

These safety matches are somewhat inconvenient because one needs a box. The latest type of match, the "bird's-eye," is a phosphorus match. The phosphorus is not, however, mixed with the entire head, but concentrated on the tip (the "eye"). The rest of the head is of the safety match variety. Such a match cannot easily be ignited by rubbing except on the tip, and this is not likely to happen.

Handling white phosphorus is a dangerous trade, causing an unsightly disease, called necrosis. Most of our matches now contain phosphorus sulphide, P_4S_3 , which is non-poisonous but still sufficiently combustible. Percussion caps used in old-style rifles contained a matchhead mixture: $KClO_3 + Sb_2S_3 + S$. This mixture can be fired by a blow and sets off the main charge.

295. Explosions of liquid combustibles. We have seen that solid combustibles may become explosive when finely divided and mixed with the supporter of combustion. This is just as true of liquids, of course, with the difference that it is much easier to break up a liquid into small particles. Some liquids, like alcohol, ether, gasoline, and carbon disulphide, vaporize readily. And what is a vapor but the liquid in gaseous form? The molecules have mingled with those of the air, and we have an explosive at once. That is why it is dangerous to light matches in a garage or wherever gasoline may be exposed to air. To assist in the rapid formation of small particles, the liquid is sometimes sprayed through a needle valve (like that used on the garden hose). This is done with gasoline in the gasoline stove, the plumber's torch, gasoline lamps, and the gasoline engine. The power furnished by the gas engine is due to a series of explosions. The mixture of air and fuel (either gas or vapor) furnished by the carburetor is drawn into the cylinder, compressed, and ignited by an electric spark across the

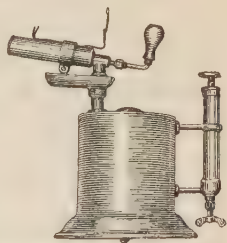


FIG. 209. Plumber's gasoline torch.

terminals of the spark plug. It is found that explosive mixtures are more sensitive when compressed, probably because the molecules are brought more

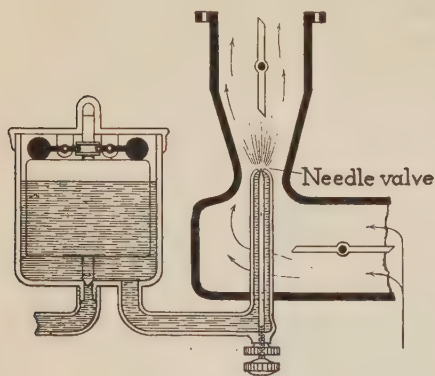


FIG. 210. Needle valve of a carburetor.

closely together. In fact, if there is no compression of the gases, the explosion is slow and therefore much less efficient. One must be exceedingly careful when handling gasoline or other combustible liquid. There are cases on record of explosions caused by cleaning a pair of gloves in gasoline.

It is even possible that the rubbing during dry cleaning may produce a spark sufficient to ignite the explosive mixture. It is safer, therefore, to use carbon tetrachloride, a non-combustible, in removing stains.

When the combustible liquid is not very volatile, the danger is very much less. Kerosene, for instance, cannot be ignited by a match held above the liquid. Just as soon, however, as the surface of contact between combustible and supporter are sufficiently increased, then the possibility of increased combustion, and even explosion, is brought nearer. For instance, kerosene on a wick or piece of cloth or even on a rough wood surface burns readily at the touch of a lighted match, and atomized kerosene will explode.

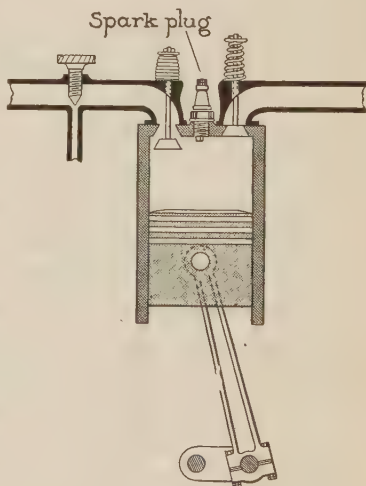
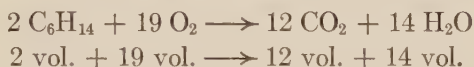


FIG. 211. Automobile cylinder.

To secure efficiency in stoves and engines, using non-volatile oils, it is, therefore, usually necessary to preheat and atomize the fuel. That explains why it is difficult to start a gasoline engine in winter. Gasoline in winter is not so volatile and requires a little heat to break it up into small particles, so necessary for the explosive mixture.

296. What makes an auto go? Let us make a rough calculation of the pressure produced by a gasoline explosion. Assume the formula of gasoline to be C_6H_{14} . Its complete combustion would be



The volume of the products of combustion is then, roughly, $1\frac{1}{2}$ times the original volume. If now the temperature rises to $1700^\circ C$. (about $6 \times 273^\circ$), the volume will again increase sixfold or to about ten times the original ($6 \times 1\frac{1}{2} = 9$). The original pressure — atmospheric — of 15 lbs. per square inch becomes, then, 150 lbs. per square inch, or about the working pressure of a steam engine.

297. Explosions of gases. Any one who has understood the foregoing paragraphs can readily guess that gases produce explosions much more often than liquids or solids. And the reason is not far to seek. Gases exist in the molecular state. The particles are already separate from each other, and of the smallest possible size. Therefore, an intimate mixture with the supporter is easily attained. In lighting a gas



FIG. 212. Burner of a gas stove.

burner one always gets a slight explosion, because the gas is first allowed to mix with air before being lighted; but the explosion is only small because of the small quantity of gas permitted to escape. Sometimes, however, fatal accidents have been the result. These have always been caused by allowing the gas to mix with air for a much longer time than necessary, before ap-

plying the lighted match. The result is an explosion of a much greater volume of gas. Persons have been known to be blown across a room, and in some cases killed, as a consequence of their ignorance or carelessness. Gas ovens are usually the chief source of trouble, especially where they are near the floor; for many people, unwilling to bend down to see that the burner is lighted, apply a taper and take the rest for granted. Meanwhile the gas may be escaping and mixing with the air. Several minutes afterward, a smell of gas gives warning that the burner is not lighted; and here lies the danger! We now have a fairly large volume of gas mixed with air, and all it needs is the kindling temperature to explode. The proper procedure in such a case is, not

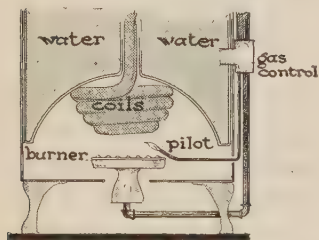


FIG. 213. Automatic hot water heater.

to light a match, but to shut off the gas and open the oven door and windows for a few minutes. This will dilute the explosive mixture with air until it is no longer explosive. Large oven burners are usually provided with a "pilot," or small burner with blue or yellow flame. This is first lighted, then the gas is turned into the large burner and lighted

from the pilot, after which the pilot may be turned off. Large automatic hot-water heaters are provided with these pilots, always burning; so that when the water is turned on, its flow turns on the gas, which is kindled by the pilot light. Once in a while, such a pilot should be inspected, for if it becomes clogged, and the flame extinguished, turning on the water will turn on the gas, but it will not take fire. And here arises another source of danger. These large hot-water heaters are usually in the cellar or some other dark place, and when hot water fails to flow from the faucet, one must not enter the place with a lighted match, because gas has been flowing and mixing with air, making an explosive mixture.

Gassy coal gives off combustible gases, particularly marsh gas (CH_4), which, mixing with air, sometimes produce the proportions necessary for explosive mixtures. Such mixtures explode at about 650°C ., if touched by a flame of one tenth second duration. It was for such mines that Sir Humphry Davy devised the miner's safety lamp. This is simply an ordinary oil light, completely surrounded by copper wire gauze. If such a lamp is placed in an explosive mixture of gases, a series of small explosions may take place inside the lamp, but the outside gases cannot reach their kindling temperature; for the hot products of combustion, produced in the lamp, must pass through the wire gauze to get out, and



FIG. 214.
The Davy
safety lamp.

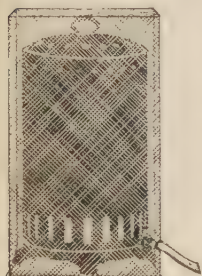


FIG. 215. Gas heater
in a garage.

in so doing, come into contact with the copper wires, which absorb the heat. The wire gauze may be looked upon as a heat filter; a hot gas passed through comes out a cool gas, but the wires never get as hot as 650°C .; in fact, they feel just warm to the hand. (The English miners, delighted with this invention, presented Davy with a sum of money, which he turned over to the Royal Society, thus establishing the Davy Medal.) In winter, some garages are heated

by gas, in which case it is safest to inclose the gas heater in a cage made of copper gauze.

298. Self-lighters. One form of self-lighter contains palladium wire, which adsorbs gaseous alcohol or gasoline, causing compression. This compression gives rise to heat and sets fire to the alcohol or gasoline vapor. Other devices set fire to the explosive mixture of vapor and air by means of sparks.

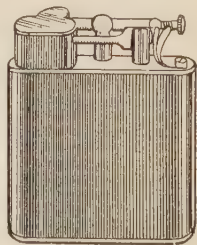


FIG. 216. A self-
lighter.

299. Dust explosions. We have seen that, in order to explode, a solid combustible must be finely divided and thoroughly mixed

with the supporter of combustion, and any solid that can burn can become explosive. In coal mines, therefore, when the coal is blasted, a great deal of dust is sometimes produced, and this dust floating in the air becomes an explosive. This state of affairs is minimized by using what are called "permissible explosives," which do not crush and shatter the coal and whose flames are short ($\frac{1}{5000}$ sec.) and not above 2000° C. Explosions in flour mills, chocolate factories, and sardine canneries have resulted from similar causes. In fact, any combustible substance, finely divided and thoroughly mixed with oxygen, becomes an explosive.

Air is not the only source of oxygen, however. Many solids, like KNO_3 and KClO_3 , readily give up their oxygen, so that if a solid combustible, like charcoal, were powdered and mixed with KNO_3 , we should have an explosive, resembling black gunpowder. Paper dipped in a saturated solution of potassium nitrate and dried, burns very rapidly, almost explosively. This is tinder paper. KClO_3 gives up its oxygen even more easily than KNO_3 ; for this reason, mixtures containing KClO_3 , explode on being struck or rubbed, and are, consequently, extremely dangerous to handle.

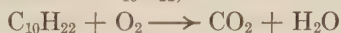
SUPPLEMENTARY READING

- R. M. Yerkes, *New World of Science* (The Century Co., 1920), chaps. viii and ix.
Geoffrey Martin, *Modern Chemistry and Its Wonders* (D. Van Nostrand Co.), chap. iii.
H. E. Howe, *Chemistry in Industry* (Chemical Foundation, 1924), chap. xii (matches).

QUESTIONS

1. What two kinds of material are necessary to make an explosive? Give examples of each kind.
2. Explain how a lump of sulphur may be made into an explosive.
3. Why do liquids like gasoline and carbon disulphide (CS_2) explode easily?
4. Why does Sb_2S_3 make a less violent explosive with KClO_3 than the element sulphur?

5. Suggest a mixture, not mentioned in the text, which might be used in match heads.
6. Why is powdered glass used on safety-match boxes?
7. Why are safety matches "safe"?
8. Mention the constituents of the head of a match, and explain the need for each substance.
9. Explain the purpose of the wick in an oil lamp.
10. Explain how the explosion of the gasoline is brought about in the automobile engine.
11. Why is ether sometimes added to gasoline?
12. Explain the danger of cleaning things in gasoline.
13. Why are gases most easily explosive?
14. Outline the proper way to light a large gas burner.
15. Make a diagram of a gas stove in a garage, properly protected (use principle of the Davy safety lamp). Explain the safety device.
16. Why are explosions frequent in chemical experimentation?
17. Describe how one type of self-lighter works.
18. Why does not coal dust burn as well as lump coal, in a stove?
19. Explain the need for regulation of a carburetor in a gas engine.
20. Explain how the thistle tube may act as a safety device.
21. In dealing with hydrogen, explosions are very frequent. What gives rise to such explosions and how may they be prevented?
22. What kind of material is most likely to explode? Why?
23. How much air would be needed to burn completely 10 liters of marsh gas, CH_4 ?
24. In the combustion of $\text{C}_{10}\text{H}_{22}$,



Complete the equation and compare the volume of the materials used with that of the products.

25. Explain why a nitrate may be used in an explosive mixture.
26. Name one compound of sulphur formed during the explosion of gunpowder.
27. One of the products of the combustion of gunpowder is nitrogen. Why is there not produced a compound of nitrogen?
28. In the explosion of dynamite, carbon monoxide is formed. What might be added to lessen the amount of carbon monoxide?
29. A solution of white phosphorus in carbon disulphide takes fire spontaneously when exposed to air. Explain.

OPTIONAL QUESTIONS

30. Explain the difference between a fuel and an explosive.
31. In the equation $\text{CH}_4 + 2\text{O}_2 \longrightarrow \text{CO}_2 + 2\text{H}_2\text{O}$, show that oxygen acts as an oxidizing agent while CH_4 acts as a reducing agent.

CHAPTER XXIV

SODIUM COMPOUNDS

300. If you will look at the second column of the periodic table on page 51, you will find the elements lithium, sodium, potassium, rubidium, and caesium. They resemble one another in many ways. For example, they are all monovalent, which means, judging from our chapter on the atom (chap. vii), that they are all very active, chemically speaking. So active, indeed, are they, that they easily oxidize when exposed to air; and they are, therefore, kept in an inert liquid, like petroleum. In contact with water they react violently, liberating hydrogen and forming the corresponding base.

In this chapter we shall confine ourselves to sodium and its compounds, for sodium is the commonest of these elements; but very much of what is said about sodium and sodium compounds applies to potassium and its compounds, and to the other alkali metals — caesium, rubidium, etc. — and their compounds.

301. Sodium itself, which is used very extensively in the preparation of various dyes and drugs, and also in the preparation of sodium peroxide, Na_2O_2 — which, in turn, may be used to prepare hydrogen peroxide and oxygen — is made by a method first used by Davy in 1807, namely, electrolyzing fused sodium hydroxide. By “fused sodium hydroxide” we mean sodium hydroxide in the molten condition. A current of electricity will not pass through the solid material, but when it is first heated until it melts, the current will pass through, the fused sodium hydroxide becoming a conductor of electricity although no water is present. Davy’s method of obtaining sodium (and, in a similar way, potassium can

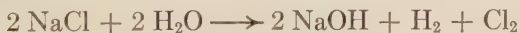
also be obtained) is of great historical interest ; for until his time no one had succeeded in actually isolating any of the alkali metals. Indeed, it was generally supposed that such substances as sodium hydroxide, NaOH, and potassium hydroxide, KOH, were elements !

A very interesting development in the use of sodium is as a catalyzer in the manufacture of artificial rubber.

302. Sodium chloride, or common table *salt*, may be made directly by burning sodium in chlorine. This remarkable reaction has already been referred to elsewhere (chap. viii). On a large scale, it is obtained from the ocean (which contains nearly three per cent of it), or from salt lakes, or from salt beds. Near Syracuse, New York, enormous salt deposits are found. By sinking shafts and pumping water into them, the salt is brought to the surface in the form of a concentrated solution. The water is driven off by heating in vacuum pans (under reduced pressure) and the product is further purified by redissolving it in water and allowing the salt to crystallize.

As a food, salt is absolutely indispensable. Since it is very cheap, salt is the source, also, of most of the other sodium compounds, such as the hydroxide, the carbonate, and the bicarbonate. From it hydrochloric acid and chlorine are made. In the making of ice cream, in refrigeration, as a food preservative, in the manufacture of soap, and in the process of tanning, salt also finds important uses.

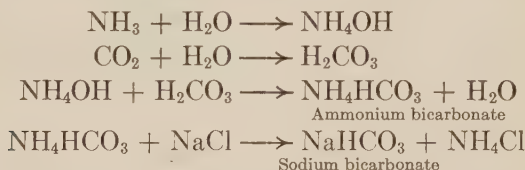
303. Sodium hydroxide, soda, or *caustic soda* ("caustic" because the substance acts on the skin) is made by the electrolysis of brine, a process which will be described in a later chapter :



It is one of our commonest of bases, and one of its most important uses is in the manufacture of soap. But caustic soda is also used in the refining of petroleum, in the manufacture of artificial silk, in the preparation of pulp paper, and as a cleansing agent (the "lye" of commerce).

Sodium hydroxide is one of the strongest bases we have; and this, as has been shown (chap. x), is due to its high dissociation into hydroxyl ions when in solution. It attacks the skin, and great care should, therefore, be taken in handling solutions of lye or other preparations containing sodium hydroxide.

304. Sodium carbonate, commonly called *washing soda*, and also known as soda ash and sal soda, is made very largely by the *Solvay* process, which, in principle, is not difficult to understand. Incidentally, this same process also enables us to get sodium bicarbonate (bicarbonate of soda or baking soda). To prepare these sodium compounds, a solution of salt (brine) is saturated with ammonia, and carbon dioxide is forced in. The reactions which take place may be presented thus:



The solutions are so concentrated that the sodium bicarbonate, which, like sodium salts, in general, is soluble in water, precipitates out in the presence of so limited a quantity of water. Since sodium bicarbonate itself is an important compound, only portions of it are used to manufacture the sodium carbonate. This is done by simply heating the bicarbonate:



Sodium carbonate, when crystalline, has the formula



and is commonly spoken of as washing soda or sal soda. On standing it loses its water and becomes anhydrous (efflorescent); it is now known as soda ash or plain soda. It is used in very large quantities in the manufacture of glass, in the softening of water, in the preparation of washing powders (owing to the fact that it

hydrolyzes and forms hydroxyl ions), and in the paper and pulp industry. In the household, the washing soda is often used to clean utensils, and its cheapness makes it serviceable in operations where acid solutions have to be neutralized. For example, solutions containing sulphuric acid may be treated with sodium carbonate to get rid of the acid:



Sodium bicarbonate, or baking soda, is used, as its name implies, in baking, as a substitute for yeast. The baking powders on the market contain sodium bicarbonate mixed with some substance which when dissolved in water forms a weak and harmless acid. Such a substance may be cream of tartar or alum, or calcium acid phosphate, which, when mixed with sodium bicarbonate, and when water is added, causes the liberation of carbon dioxide.

For baking purposes the sodium bicarbonate is to be preferred to the carbonate for two reasons: in the first place it is less alkaline and, therefore, better adapted as a food; and in the second place, weight for weight, it gives more carbon dioxide than the normal carbonate. Its weak alkaline properties make the bicarbonate of service also as a medicine — in neutralizing excess acidity in the stomach, for example.

305. By-products. The Solvay process has often been referred to as a model industrial process because so little of the various substances produced are wasted. As the chemist puts it, the “by-products” are properly made use of; and it is always the great ambition of the chemist and the manufacturer to find good uses for his by-products. Consider, for example, the series of reactions which have just been given. You will notice that besides obtaining sodium bicarbonate, we also get ammonium chloride. Should this be thrown away? No, indeed! The manufacturer takes this ammonium chloride, adds slaked lime, distils the mixture, and gets ammonia, which, of course, can be used over again for the manufacture of a fresh batch of sodium bicarbonate.

You will also notice that in the preparation of sodium carbonate from sodium bicarbonate, carbon dioxide is evolved. This, also, is used over again.

306. Sodium sulphate, $\text{Na}_2\text{SO}_4 \cdot 10 \text{H}_2\text{O}$, or *Glauber's Salt*, is used in medicine and in glass making.

307. Sodium nitrate, or **Chile saltpeter**, has already been referred to in the manufacture of fertilizers (chap. xx) and in the making of nitric acid (chap. xviii). It is also used in making potassium nitrate, KNO_3 (saltpeter or niter), a constituent of gunpowder and fireworks, and sodium nitrite, NaNO_2 , which, in turn, is used extensively in the manufacture of dyes.

308. Sodium hypochlorite (Javelle water), NaClO , which can be made by passing chlorine into sodium hydroxide, is a bleaching agent. During the late war, Dr. Dakin used it in several modifications as a disinfectant and it became known as *Dakin's solution*.

309. Borax, $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10 \text{H}_2\text{O}$, is a sodium salt which in solution is alkaline in reaction and, therefore, finds a use as an ingredient of washing powders. When treated with an acid, we get *boric acid*, H_3BO_3 , which is a common antiseptic.

310. Sodium thiosulphate, $\text{Na}_2\text{S}_2\text{O}_3$, is the photographer's "hypo." *Sodium benzoate* is used as a preservative.

QUESTIONS

1. Why is not sodium chloride manufactured directly from sodium and chlorine?

2. Starting with salt, show how we can prepare from it (a) hydrochloric acid, (b) sodium hydroxide, (c) sodium carbonate, (d) sodium bicarbonate.

3. The value of by-products often determines the success of a particular industry. Mention industries other than the Solvay process in which by-products are made use of.

4. Give the chemical names for (a) salt, (b) washing soda, (c) baking soda, (d) Javelle water.

5. Starting with sodium, show, by means of equations, how you could prepare (a) sodium hydroxide, (b) sodium chloride, (c) sodium sulphate.

6. Show by means of equations and the use of molecular weights

that, weight for weight, more carbon dioxide is obtained from sodium bicarbonate than from sodium carbonate, when each is treated with an acid.

7. Suggest chemical methods for distinguishing (a) sodium from sodium hydroxide, (b) sodium carbonate from sodium bicarbonate.

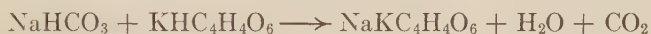
8. Explain in terms of ionic equations why a solution of sodium carbonate turns red litmus blue.

9. Suggest a method of preparing sodium sulphate (a) by simple replacement, (b) by neutralization, (c) by volatility.

10. How much (a) sodium carbonate and (b) sodium bicarbonate could be obtained from ten kilograms of sodium chloride by the use of the Solvay process?

11. How many liters of carbon dioxide and ammonia would be needed in question 10?

12. Baking soda and cream of tartar in baking powders react as follows:



How much baking soda is required for 47 grams of cream of tartar? What volume of carbon dioxide will this set free?

13. Ammonia is used in the manufacture of ice and in the manufacture of sodium carbonate. Show how the use of ammonia differs in these two cases.

14. Write equations to represent the reactions of (a) sodium hydroxide, (b) sodium carbonate, and (c) sodium bicarbonate with each of the following acids: nitric acid, sulphuric acid, hydrochloric acid.

15. What is hypo, borax, niter, Chile saltpeter, lye, soda?

16. Draw diagrams to show the electron structures of the sodium atom and the sodium ion. Now explain why sodium chloride does not react with water as does metallic sodium.

17. Using ionic equations, explain why sodium carbonate is a stronger alkali than sodium bicarbonate.

18. Why is soda used to wash woodwork before painting?

19. In the equation $2\text{Na} + \text{Cl}_2 \longrightarrow 2\text{NaCl}$, which is the oxidizing agent and which is the reducing agent? Explain.

20. Write the equation for the preparation of nitric acid from potassium nitrate. In actual practice, why is sodium nitrate preferred to potassium nitrate?

21. How could the percentage of wool be determined in a sample of cloth?

22. Given that the symbol for lithium as Li, write the formulas for the oxide, the hydroxide, the chloride, the sulphate, the nitrate, the carbonate, and the bicarbonate. How would solutions of these substances act towards litmus? Explain.

23. Why are LiCl , NaBr , and KI neutral to litmus?
24. Determine the percentage of water in crystalline sodium carbonate.
25. How does the percentage of nitrogen in NaNO_3 compare with that in CaCN_2 (calcium cyanamide)?

OPTIONAL QUESTIONS

26. A certain solution contains 0.049 gram of sulphuric acid in each cubic centimeter. How many cubic centimeters (c.c.) of this solution would be needed to neutralize 100 c.c. of a solution of sodium hydroxide which contains 0.04 gram per cubic centimeter?
27. Why was it found difficult to isolate sodium?
28. The per cent of salt in the ocean is constantly increasing. Explain.
29. Why does a hunter lie in wait at a salt lick?
30. Of what value is sodium chloride to the body?
31. Keeping in mind the law of mass action, explain why salt is precipitated from a concentrated solution by the addition of HCl .
32. Write equations to show how you would prepare (a) KOH , (b) KHCO_3 , and (c) K_2CO_3 .
33. Why is KCl used as a fertilizer?
34. Remembering the action of HClO as an oxidizing agent, explain how NaClO may also act as an oxidizing agent.

CHAPTER XXV

CALCIUM COMPOUNDS

311. Calcium is a necessary plant food, it is also a very necessary animal food. The teeth and bones of animals consist very largely of calcium phosphate, and it is, therefore, most important that the growing child should be supplied with foods rich in calcium. As this chapter will show, a number of compounds of calcium are of extreme importance in the industries. Lime, mortar, cement, and plaster — to mention but a few substances — are all calcium compounds. The source of these various calcium products is either calcium carbonate (which is found as deposits in enormous quantities as marble, limestone, and chalk) or calcium sulphate (which is found in the form of gypsum).

312. Limestone is the commonest form of calcium carbonate, CaCO_3 . There are many varieties of calcium carbonate in nature (coral, chalk, coquina, etc.). In many of them, their origin from shells is evident. Many marine animals protect themselves with shells made of calcium carbonate, which is undoubtedly derived from the calcium bicarbonate in the ocean. The accumulation of such shells ultimately forms beds of limestone. The chief use of limestone is in the manufacture of quicklime; but it is also used in very con-



FIG. 217. Marble, a form of calcium carbonate.

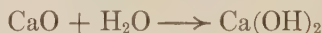
siderable quantities in the manufacture of iron and other metals, as stone for concrete mixtures and roads, and as building stone. Chalk is a form of calcium carbonate.

313. Quicklime or calcium oxide, CaO , is made by strongly heating limestone :



Since the reaction is reversible, it is necessary that the carbon dioxide be removed as fast as it is formed ; this is done by means of a

forced draft. The calcium oxide so produced is called quicklime, though it is sometimes also referred to as lime. Water combines rapidly with quicklime, forming calcium hydroxide :



Very much heat is evolved in the reaction, as can be noticed by the actual transformation of some of the water into steam. The calcium hydroxide so formed, known as *slaked lime*, is still a solid. It is slightly soluble in water, forming a solution of calcium hydroxide which is commonly known as *limewater*, and which finds extensive

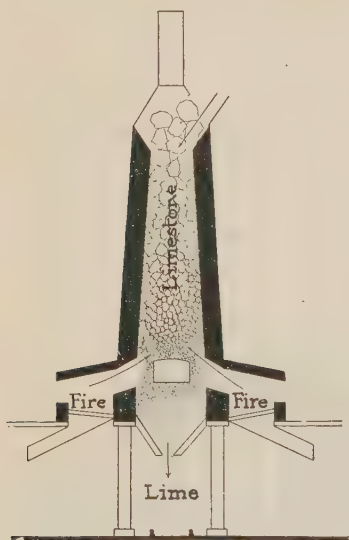


FIG. 218. A lime kiln.

use as a cheap base. Limewater mixed with copper sulphate forms what is known as *Bordeaux Mixture*, which is used as a spray to remove fungi from plants. The *milk of lime* supplied by the druggist is merely limewater containing some undissolved calcium hydroxide.

314. Mortar, so extensively employed in the building industry, is made from slaked lime by the addition of sand and water. The paste so formed will "harden" after it has been exposed to the air

for some time. This is due, in part at least, to the union of the carbon dioxide (found in the air) with the calcium hydroxide in the mortar:



The calcium carbonate so formed is hard, and, mixed with the sand, helps in keeping the brick or stone in place. For the hardening of mortar (the "setting" of mortar, as it is often called), air is necessary; in the absence of air, mortar will not set. That explains why in the building of foundations for houses, bridges, etc., where the surroundings are moist, mortar cannot be used; here cement is substituted.

Besides being used in the manufacture of mortar, lime finds a number of other important uses. It is a necessary substance in the manufacture of bleaching powder, calcium carbide, calcium bisulphite (used in paper-making), and glass. It is used in the purification of water and in removing hair from hides in tanning.



FIG. 220. Cement sidewalk.

commonly known as *Portland cement*, is to heat limestone with material containing oxides of aluminum and silica, such as clay, slate rock, or blast-furnace slag. The product consists of silicates and



FIG. 219. Making mortar.

315. Cement. When limestone containing clay is heated, a mortar is obtained which can "set" under water; it is known as *hydraulic mortar*, or, still more commonly, *cement*. The present method of making cement, com-

aluminates of calcium. As has already been stated, cement hardens when mixed with water.

316. Concrete, a most important building material, is made from cement by mixing it with sand and broken stone. Roadbeds, aqueducts, and the foundations of many buildings are made of this material. *Reinforced concrete* consists of concrete with iron rods imbedded in it. This strengthens the mass.

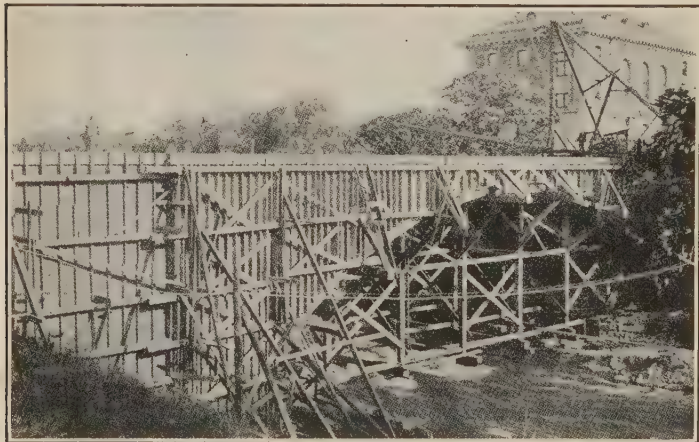
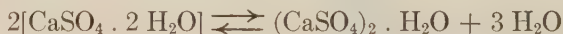


FIG. 221. Reinforced concrete.

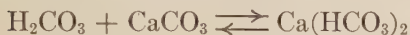
317. Gypsum, $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$, is a natural mineral found in many parts of the United States, and is the chief constituent of blackboard "chalk." When heated, it loses part of its water, yielding a product known as *plaster of Paris*:



Plaster of Paris, in turn, when moistened with water, sets in a short time to a hard mass resembling in composition that of gypsum, so that the above reaction is a reversible one. Since this plaster of Paris, in setting, expands somewhat, it is used very extensively in the making of molds. It is also used for casts and for surgical bandages when it is desired to keep parts of the injured body in a

fixed position. Mixed with slaked lime, it becomes the "plaster" commonly used for walls and ceilings.

318. Water softening. The waters we drink and use for industrial purposes contain varying quantities of salts of calcium (and magnesium). We find quite often considerable quantities of calcium bicarbonate, $\text{Ca}(\text{HCO}_3)_2$, in solution. This salt arises from the interaction of carbonic acid (found in water) and calcium carbonate (found in the region):



This solution of the limestone accounts for the "caves" found in limestone regions, such as the Mammoth Cave in Kentucky and the Luray Caverns in Virginia.

The calcium carbonate is but slightly soluble in water, whereas the calcium bicarbonate is very much more soluble; hence considerable quantities of the latter may be found dissolved in water. By boiling such a water, the above reaction is reversed,

the calcium bicarbonate being decomposed into calcium carbonate, which precipitates out, and the carbonic acid breaks up, liberating carbon dioxide. Iron in the form of $\text{Fe}(\text{HCO}_3)_2$ may also be present; if it is, and if the water is boiled, it also decomposes, the iron being precipitated as FeCO_3 . It is these

precipitates which often give rise to deposits or "fur" in the kettle and to "scale" in the boiler. Waters containing bicarbonates of calcium or magnesium are called *temporary hard waters*.

Very often the sulphates of calcium and magnesium are found in water, which cannot, then, be softened by boiling. Such waters are therefore called *permanent hard waters*.

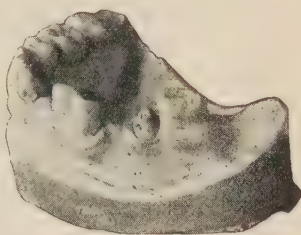


FIG. 222. Plaster of Paris is used in dentistry.

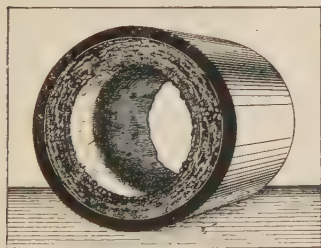
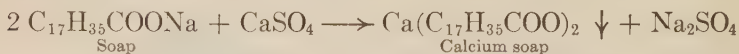


FIG. 223. Boiler scale.

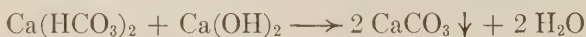
Now it so happens that waters containing calcium or magnesium salts in solution will react with soap solutions, giving slimy precipitates instead of the usual lather. This is due to the following type of reaction :



The calcium soap, unlike the ordinary soap (which is a sodium or potassium compound), is insoluble in water, and quite valueless for washing purposes. This, therefore, means that much soap will be wasted before any lather can be obtained. Indeed, a lather will not be obtained until all of the calcium (or magnesium) ions are first precipitated out of solution.

To remedy this defect, to prevent soap wastage and staining of fabrics (due to the deposit on the fabric of the insoluble soap), various methods of *water softening* have been introduced. These water-softening processes are not only necessary for washing purposes, whether in a laundry or in the household, but they are of great importance in preventing the formation of crusts or scales in boilers ; for when the water in a boiler is heated, the calcium and magnesium salts gradually deposit on the bottom of the boiler, making it increasingly difficult to heat the boiler properly and economically.

With a temporary hard water, the water may be softened by boiling, since this precipitates out the salts of magnesium and calcium. On a large scale, however, this is not so cheap an operation as treating the water with slaked lime :



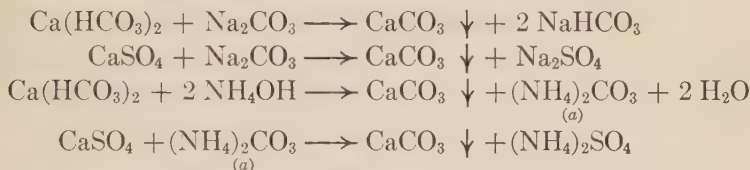
the precipitate of calcium carbonate so formed being removed by filtering.

Water containing calcium or magnesium sulphate (permanent hard water) cannot be softened in this way. Such a water can be softened by adding sodium carbonate :

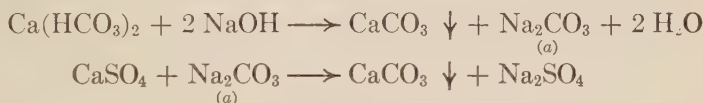


Quite often the water is both temporarily and permanently hard; in this case limewater is first added to precipitate the bicarbonates; and then soda ash (crude sodium carbonate) to precipitate the sulphates.

In homes, on a small scale, washing soda ($\text{Na}_2\text{CO}_3 \cdot 10 \text{H}_2\text{O}$) and "household ammonia" (NH_4OH) are often used to soften water. The reactions may be represented as follows:



Sometimes borax ($\text{Na}_2\text{B}_4\text{O}_7 \cdot 10 \text{H}_2\text{O}$) is also used. Here its softening properties are due to the fact that in water it is hydrolyzed, yielding, among other things, sodium hydroxide, which acts as follows:



You will notice that these various softening processes depend upon the change of soluble into insoluble calcium compounds.

Large quantities of water are now softened by the *permutit* process. Permutit itself is a complex sodium silicate made artificially. By passing the water through beds of permutit, the sodium in the permutit is replaced by the calcium of the hard water. Representing, for convenience, the permutit as NaX , we may write the equation:



By treating the CaX_2 with a salt solution, the original permutit is recovered:

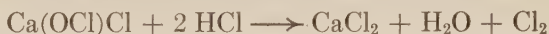


and thereby can be used over again.

319. Bleaching powder, often, and incorrectly, called chloride of lime, is, as its name implies, a substance used in bleaching. It is made by combining chlorine with slaked lime:



Acids react with bleaching powder, liberating chlorine:



and this reaction explains its use as a bleaching agent (see chlorine, chap. xiii). However, much of the bleaching powder of commerce is now being displaced by liquid chlorine. Bleaching powder, like chlorine, is a disinfectant, and for the same reasons.

Several other important calcium salts, such as calcium phosphate, calcium carbide, and calcium cyanamide are discussed elsewhere.

SUPPLEMENTARY READING

"Lime Symposium," *Industrial and Engineering Chemistry*, vol. XIX (1927), pp. 550-605.

G. A. Rankin, "Portland Cement," *Chemistry and Industry*, vol. II, p. 261.

QUESTIONS

- Starting with calcium carbonate, show how you could get (a) calcium oxide, (b) calcium hydroxide, (c) bleaching powder.
- What are some names commonly used to represent the substances mentioned in (1)?
- What is mortar and how does it "set"?
- Compare the hardening of mortar, cement, and plaster of Paris.
- Show, with the help of equations, how you would soften a water containing calcium bicarbonate and magnesium sulphate dissolved in it.
- What type of water is necessary (a) for drinking purposes, (b) for washing, (c) for use in a boiler?
- When limewater is exposed to the air, it turns milky. Explain.
- Wood ashes have been used to soften water. Can you offer an explanation?
- Describe the principle of the permutit process.
- Suggest a method for distinguishing (a) calcium carbonate from calcium oxide, (b) calcium hydroxide from bleaching powder.

11. Chlorine gas is used in the manufacture of bleaching powder, which, when used, sets free chlorine. Why is not the original chlorine used rather than the bleaching powder?

12. Mention four calcium compounds and give an important use of each.

13. Write equations to represent the reactions of (a) calcium oxide and (b) calcium hydroxide, with each of the following acids: nitric acid, hydrochloric acid, sulphuric acid.

14. What weight of calcium oxide could be obtained from 200 grams of pure calcium carbonate? What volume of carbon dioxide measured at standard conditions would be liberated in the reaction?

15. Limewater is often added to cow's milk when used by infants. Give two reasons why limewater is used in place of sodium hydroxide.

16. Why does lime exposed to the air become useless?

17. What is concrete, plaster, cement, boiler scale?

18. $\text{Ca(OH)}_2 + \text{CO}_2 \longrightarrow \text{CaCO}_3 + \text{H}_2\text{O}$. Name as many processes as you can which this equation represents.

19. Cement hardens throughout, mortar does not. Explain.

20. How has the increased use of concrete affected the building industry, maintenance of highways, building bridges and tunnels?

21. Why are foods containing calcium so important to growing children?

22. How can the setting of plaster be retarded (refer to chap. xiv)?

23. What is meant by "phosphorus" in fertilizers?

24. Write an ionic equation to show the reaction of calcium carbonate and an acid.

25. Show how you could make sodium hydroxide from calcium hydroxide. Why does the reaction go to completion?

26. How would you distinguish NaOH from Ca(OH)_2 (test for CO_2); Na_2CO_3 from CaCO_3 ?

27. Write an equation to show that quicklime is a basic anhydride.

28. Write the reaction between CaCO_3 and HNO_3 . Compare this reaction with Ca(OH)_2 and HNO_3 . Starting with CaCO_3 , write equations to show two ways of obtaining CO_2 .

29. Explain the use of an open fire to hasten the setting of "plaster."

CHAPTER XXVI

METALS

320. In ancient times, men came upon deposits of mineral by chance, and discovered how to extract metals by accident. In this industrial age, and especially with the help of our knowledge of science, expert geologists are sent out to scour the earth for material to satisfy the enormous demands of modern life. Our government and many of the states maintain geological surveys; but it is usually the business man who finances the enterprise to search for new deposits of mineral ("prospecting").

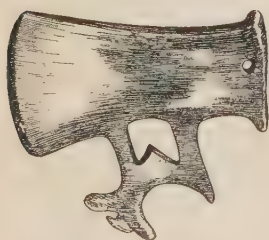


FIG. 224. Indian axe head
made of copper.

A *mineral* is any substance found in nature, not plant or animal, which has a definite chemical composition. Now, the value of a mineral deposit depends upon many factors. It must not be too remote from civilization, or the cost of transportation will become prohibitive; it must be easily mined, that is, taken out of the earth; and it must be of the proper degree of

purity. If you keep constantly in mind the idea that business means profit, then it can easily be understood why some mineral deposits are used, while others are neglected. Pyrites, or fool's gold, FeS_2 , is one of the commonest minerals of the earth, and it contains a great deal of iron; but it is not used for the production of iron, because the metal can be obtained more easily from hematite, Fe_2O_3 ; and besides, it contains too much sulphur, which makes a poor grade of iron. Pyrites is an iron mineral, but

it is not an ore of iron. *Ores* are those minerals from which we can profitably extract useful substances. Hematite is the chief ore of iron, but perhaps, after all the hematite and other present ores of iron are exhausted, we may have to use pyrites to get our iron, and then pyrites will be the ore of iron.

321. Classification of ores. Most metals are easily attacked by oxygen, carbonic acid, and sulphur, and since these occur in nature in large quantities, it is evident that those metals which are easily attacked will not be found free (or native) to any great extent. The metals which resist such chemical change, like platinum and gold, are called precious or noble, while the others are called base metals. Precious metals are, therefore, found native, but there are also deposits of native base metal, principally copper, which have not yet been attacked, because of the

absence of sulphur-bearing waters, and because the air has not yet penetrated the depths. In other words, they are "young" deposits. In time, practically all minerals become oxidized, so that a deep deposit of copper sulphide, for instance, will show as oxide or carbonate near the surface. Sulphides also change to sulphates, but since most sulphates are soluble in water, these will be carried into the ocean, so we find only those that are insoluble, like sulphates of lead, barium, calcium, and strontium. Only silicates remain practically unchanged, and most minerals are silicates. But silicates are hard to treat, or smelt, as we say; therefore,



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FIG. 225. "Panning" for gold.

most ores are not silicates, but oxides and sulphides. We may classify most ores, then, as follows :

KIND OF ORE	EXAMPLE
Native	Gold, Platinum, Silver, Copper
Oxidized ores {	oxides zincite (ZnO); hematite (Fe ₂ O ₃); bauxite (Al ₂ O ₃); cuprite (Cu ₂ O)
	carbonates smithsonite (ZnCO ₃); siderite (FeCO ₃); malachite (CuCO ₃ · Cu(OH) ₂); cerussite (PbCO ₃)
	silicates willemite (Zn ₂ SiO ₄)
	sulphates anglesite (PbSO ₄); barytes (BaSO ₄)
	Sulphides argentite (Ag ₂ S); sphalerite (ZnS); cinnabar (HgS); galena (PbS); chalcopyrite (Cu ₂ S · Fe ₂ S ₃)
Miscellaneous	sylvanite (AgAuTe ₂); horn silver (AgCl)

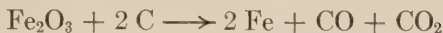
322. Treatment of native ores. Native metals, like gold and silver, require very little treatment, besides separation from their gangue, the rock in which they are found. This usually consists of crushing the ore, so as to break particles of rock from particles of metal; separation of the two as well as possible, by moving them in water; and finally fusion with a *flux*, to remove the rest of the gangue. The flux is so called because it makes the rock fluid. It is usually limestone, CaCO₃. Rocks being chiefly infusible silica and silicates, that is, acid in their reaction, require a base-forming material to combine with them; and limestone is just such a substance, and it is, furthermore, abundant and cheap. When heated, it decomposes as follows :



and this basic CaO combines with SiO₂ and the silicates of other metals to form a mixed silicate, called *slag*, which melts as easily as glass.

323. Treatment of oxidized ores. Any oxidized ore, if pure enough, may be reduced directly, some form of carbon being used :

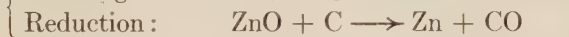
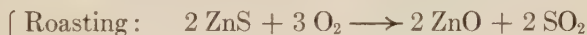
coke or charcoal, usually. Of course a flux must be used to take care of the gangue. Sample reactions are given for each of the oxidized ores :



We do not usually find pure oxidized ores except close to the surface. They are found mixed with sulphides, so that they are frequently thrown in with, and treated as, sulphides, since it does not pay to attempt separate treatment.

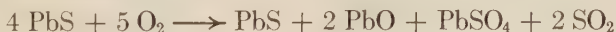
By far the greater quantity of the ores of metals, other than iron, are sulphides.

324. Treatment of sulphide ores. With these ores, two operations are necessary. A preliminary roasting (heating with air) drives off most of the sulphur as SO_2 , leaving the metal as oxide. This is then treated like an oxidized ore ; that is, it is reduced with carbon. Examples are given for several metals :

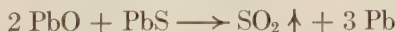


Formerly, lead sulphide (galena), the chief ore of lead, was so treated ; but the modern process consists of partial roasting, and reduction by the remaining PbS , as follows :

Partial roasting :



Air is then shut off and the temperature raised ; then,



The chief ore of copper, chalcopyrite, is treated in a very complicated fashion, due chiefly to the desire to collect and save the so-

called "values" (gold and silver) whose presence is often the factor that makes the process of extraction profitable. After the copper is separated and refined, electrolytically, there remains in the bottom of the tank a mud or "sludge," as it is called, which contains, among other things, gold and silver. This is dried and fused with lead, which extracts the gold and silver. This "base bullion" is then heated in air, and the lead oxidized, leaving a mixture of gold and silver. The silver is dissolved by nitric acid, leaving the gold. The silver in the silver nitrate is then recovered by replacement with copper.



Sulphides of the precious metals may be treated in one step: oxidation, because the metal, when freed, does not readily combine with oxygen. This is also true of mercury, which is therefore extracted as follows:



325. Electrolytic processes of extraction. The high cost of labor and fuel is bringing with it the demand for processes requiring less fuel and labor. Where electric power is cheaply produced, the electrolytic processes are coming into vogue. They are simple, require little labor, and produce better products. Aluminum is

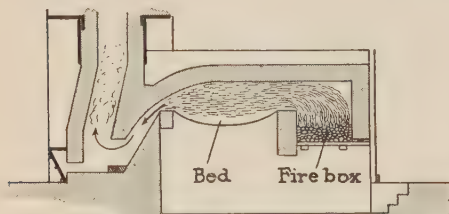


FIG. 226. Reverberatory furnace.

produced in this way. Sodium, potassium, and calcium are produced by electrolysis, zinc and lead are sometimes so extracted, and copper, except from native ores, is refined electrolytically.

326. Furnaces. For fire metallurgy, various types of furnaces are used, and it is perhaps worth while making a very brief study of them. As we

have seen, there are two essential chemical operations carried out in furnaces: oxidation and reduction. Corresponding to these, we have the following:

TYPE OF FURNACE	CHEMICAL ATMOSPHERE
Reverberatory	oxidizing
Blast furnace	reducing
Bessemer converter	oxidizing

In the reverberatory furnace, the fire box is separate, preventing contamination. For roasting, a reverberatory furnace is used, usually of a special type, with mechanism to stir the ore, so that it may become thoroughly oxidized. The "open hearth" is a reverberatory furnace so arranged that the metal can be poured.

In the blast furnace the ore and fuel are mixed, so that the product is impure; but large quantities can be handled, since the products run out at the bottom by gravity.

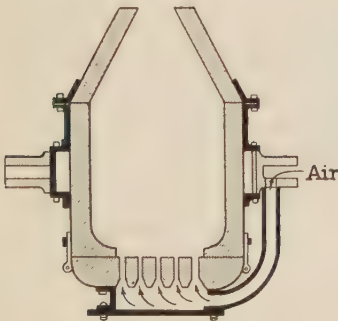


FIG. 228. Bessemer converter.

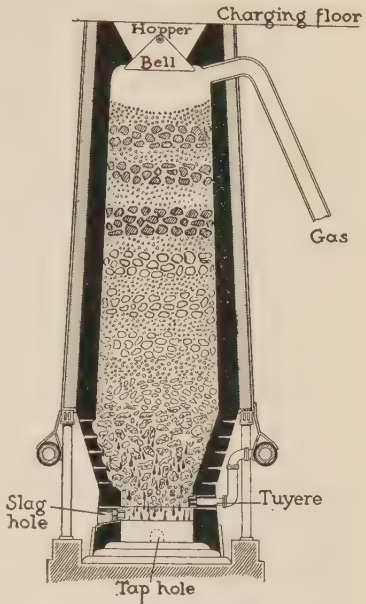


FIG. 227. A blast furnace.

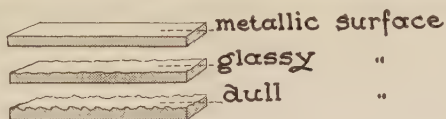
The Bessemer converter is a modified blast furnace, but since no fuel is added, the atmosphere is an oxidizing one. It is used to purify metal by burning out certain elements, like carbon, sulphur, and phosphorus; and it is these impurities which burn and produce the heat. Some of the oxides formed pass off as gases; others are slagged.

327. Properties of metals. Many metals find very important uses in the arts, and since these uses depend upon certain properties, it is well to understand the meaning of the words which name the properties. And it is just as well to realize, right at the start, that when it is decided to use a certain metal, the chief factor which determines the use is the price. The value of gold is fixed at \$20.67 per ounce, Troy. It is the standard of values throughout most of the civilized world. Other metals and commodities vary in value from day to day, quotations being published in many newspapers, by the metal and other exchanges. Just to give some idea of the relative values, however, the following table of approximate prices is inserted :

iron	about 1 cent per pound
zinc	about 6 cents per pound
lead	about 6½ cents per pound
copper	about 16 cents per pound
aluminum	about 24 cents per pound
tin	about 50 cents per pound
nickel	about 80 cents per pound
silver	about 66 cents per ounce
gold	\$20.67 per ounce, Troy
platinum	\$93.00 per ounce, Troy

328. Luster. The most outstanding property of metals is the luster. This is more easily understood and recognized than defined in words. Every one knows the difference in appearance between a metal like silver and a substance like stone. The silver shines, it is brilliant, it reflects light. We may, in fact, classify substances, roughly, into three classes, as to luster : metallic, vitreous or glassy, and dull. A metallic substance, when polished, shows nearly total reflection of light ; a glassy substance reflects only part of the light, while a dull object scatters all the light. If we examine these materials under the microscope, we

find the metal is smooth, the dull object is very rough, and the glassy material in between, somewhat as shown in the diagram :



The high luster, then, is not inherent in metals, but is due to the tenacity, which allows very small bits of metal to be worn off, whereas when a dull substance is polished, much larger particles of the surface are torn out.

329. Color. Most metals are white, some of bluish tinge, some grayish. The only two common metals which are not white are gold and copper.

330. Tenacity or strength. This is the property which enables metals to resist blows, as well as pressure, tension, bending, and twisting. The term includes, wholly or in part, malleability, ductility, and elasticity. The "strength" of a material is measured by the number of pounds of stress per square inch which must be exerted in order to break the material. This is called the "ultimate breaking stress."

For example, a piano wire, about the diameter of a match, has a breaking stress of $1\frac{1}{2}$ tons. The following are arranged in the decreasing order of tenacity: iron, copper, aluminum, platinum, silver, zinc, gold, tin, lead.



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FIG. 229. Steel has great tenacity.

331. Malleability. The word "malleability" comes from a Latin word which means "hammer" (mallet). The malleability of a metal is measured by the thinness to which a piece can be pounded, without cracking. One ounce of gold may be beaten out so as to occupy 160 square feet of surface. Gold leaf is about $\frac{1}{300000}$ of an inch thick. The following metals are arranged in decreasing order of malleability: gold, silver, aluminum, copper, tin, lead, zinc, platinum, iron.

332. Ductility. This is measured by the thinness of the wire that can be made of a given substance. One grain of gold may be drawn into a wire 500 feet in length. The following metals are arranged in order of decreasing ductility: gold, silver, platinum, iron, copper, aluminum, zinc, tin, and lead.

333. Hardness. This is a term that is misunderstood and misused by many people. Hardness is measured by the power to scratch or wear away another material. When we say one substance is harder than another, we mean it will make a visible scratch, if the two are rubbed together. Diamond has a hardness of 10, glass 7, marble 3. These figures refer to a table of hardness, which is here given:

- 10 diamond
- 9 ruby (or carborundum or emery)
- 8 topaz
- 7 quartz or glass
- 6 feldspar
- 5 apatite (calcium phosphate)
- 4 fluor spar
- 3 calcite (limestone)
- 2 gypsum or rock salt
- 1 talc (soapstone)

It is usually a simple matter to select a material harder than a given metal, because almost any alloy of a metal is harder than the pure metal.

Emery and carborundum, with a hardness of 9, are made up into powders, which are sized, and then used as powder, or stuck on paper or cloth (emery paper) or made into whetstones or wheels, for grinding or polishing. Such substances are called *abrasives*.

334. Brittleness. This property is the opposite of tenacity and malleability. A substance which breaks on receiving a blow, or on bending, is brittle. Hard substances, such as cast iron, are usually brittle.

335. Specific gravity. Most people say a metal is "heavy," but a pound of feathers is just as heavy as a pound of lead. What they mean to say is that considering the size, or volume, the metal is heavier than the feathers; or as we say in scientific language, the specific gravity (sp. g.) is high. We use water as the standard, calling its specific gravity 1. Anything heavier than water is greater than 1, and anything lighter than water is less than 1. If a cubic centimeter of iron weighs 7 times as much as a cubic centimeter of water, we say the specific gravity of iron is 7. Metals have high specific gravity, although lithium, sodium, and potassium are lighter than water. The specific gravity of some important metals is shown in the table:

M.G.

potassium	0.86
sodium	0.97
magnesium	1.7
aluminum	2.7
iron	7.0
zinc	7.1
tin	7.3
copper	8.8
silver	10.5
lead	11.0
mercury	13.6
gold	19.3
platinum	21.0

336. Melting point. The melting point is the temperature at which a material becomes a liquid. The melting point of metals varies very much, but the alloy of a metal has usually a lower melting point than the pure metal. A table of melting points of metals and alloys is given here :

METAL OR ALLOY	CHIEF CONSTITUENTS	MELTING POINT (C.)
Tungsten	W	3500
Platinum	Pt	1770
Wrought iron	Fe	1600
Nickel	Ni	1450
Steel	Fe, C	1250
Cast iron	Fe, C	1200
Gold	Au	1100
Copper	Cu	1090
Silver	Ag	960
Bronze	Cu, Sn	920
Aluminum	Al	658
Zinc	Zn	419
Lead	Pb	327
Cadmium	Cd	320
Bismuth	Bi	265
Tin	Sn	232
Solder	Pb, Sn	190
Rose's Alloy	Bi, Pb, Sn	95
Wood's Alloy	Bi, Pb, Sn, Cd	60.5
Darcet's Alloy	Bi, Pb, Sn, Hg	45.0

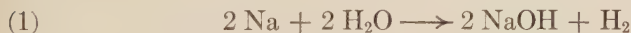
337. Heat conductivity. If one holds in a flame wires made of silver, copper, and iron, it will be noticed that the silver wire carries heat to the hand first and the iron last. This is a rough measure of their conductivity of heat. The following is a list of metals in decreasing order of heat conductivity: silver, gold, copper, mercury, aluminum, zinc, wrought iron, tin, steel, platinum, cast iron, lead, and bismuth.

338. Electrical conductivity. This term is common enough, although its opposite, resistance, is more frequently used. It will

be noticed that pure metals are much better conductors than alloys. The following metals and alloys are arranged in decreasing order of electrical conductivity: silver, copper, alloy of $\frac{1}{2}$ copper and $\frac{1}{2}$ silver, gold, aluminum, zinc, brass, wrought iron, tin, German silver, steel, platinum, lead, bronze, nickel, and antimony.

339. Other properties of metals. It is noteworthy that none of the metals dissolve, except when they undergo a chemical change. Sulphur and phosphorus dissolve in carbon disulphide, and iodine in alcohol; and, upon evaporation of the solvents, the pure elements remain, unchanged. In this sense, there is no solvent for any of the metals. Nitric acid is said to dissolve copper, but during the solution, the copper is changed to copper nitrate. Nitric acid attacks all common metals, except gold, and it is therefore used to test for gold. Hydrochloric acid attacks most common metals readily, except copper, and this is also true of sulphuric acid.

“Pickling” metals consists of immersing them in some acid (the “sour”), like hydrochloric or sulphuric, in order to clean them by dissolving the surface. This is done preparatory to plating. Most metals are easily attacked by oxygen, forming oxides, most of which are white powders. Metals replace hydrogen from its solutions:



or (1) may be written $2 \text{Na}^0 - 2 e \longrightarrow 2 \text{Na}^+$

and $2 \text{H}^+(\text{from } \text{H}_2\text{O}) + 2 e \longrightarrow \text{H}_2^0$

and (2) $\text{Zn}^0 - 2 e \longrightarrow \text{Zn}^{++}$

and $2 \text{H}^+(\text{from } 2 \text{HCl}) + 2 e \longrightarrow \text{H}_2^0$

For this reason metals are attacked by acids and water. Since we find carbonic acid and water in air, we have an ever-present source of corrosion of most metals. Of the base metals, copper is least easily attacked, because it does not replace hydrogen. Carbonic acid attacks copper, forming green basic carbonate, which protects the metal against further corrosion. The precious metals,

silver, gold, and platinum, are not attacked by oxygen, water, or carbonic acid. Sulphur compounds attack these, however, especially silver. On the latter a black coating, Ag_2S , is formed.

Metallic oxides, when soluble in water, form hydroxides, which are bases.

340. Protection of metals. Since most metals are easily attacked by oxygen, water, and carbonic acid, it is usual to protect them, by coating with some resistant material. Paint is most often used; but it soon wears off. Therefore, in many cases, the metal is plated. On iron the chief coatings used are zinc, tin, copper, nickel, and chromium. Copper or brass articles are sometimes plated with nickel, silver, or gold. White metal (nickel or its alloys), used for table silver, is usually silver plated.



FIG. 230. A tin can: iron plated with tin.

341. Uses of metals. In selecting a material for a certain use, what are going to be the deciding factors? The price and the properties. Now, only too often, the decision is based entirely on initial cost; whereas upkeep must also be considered. If a copper leader lasts fifty years, while a galvanized iron one lasts five years, it is, obviously, cheaper to use copper, even if the initial cost is twice as much.

The most useful metal is, without doubt, iron. To such an extent is this true, that metals are classified, in the trade, as ferrous (iron and steel) and non-ferrous (all others). Its low price, great strength, and abundance make it most desirable for all kinds of structural material.

342. Copper. Next in importance to iron, copper is the most useful metal. It is not nearly so strong as iron, but it has several properties which make it more useful than iron. It is an excellent conductor of heat and electricity, and for these reasons, wire to carry electric power, pipes for water-heating coils, and many similar things, are made of copper. It is not quickly corroded, and is therefore used for roofing and for sheathing ships. Copper and its

chief alloys, brass, bronze, and German silver, can be readily cast and worked with tools, and these properties make them useful for all sorts of ornamental articles. The high luster of these metals is also a desirable feature; especially brass, which can be made to resemble gold in color and luster. Much copper is made into copper sulphate, for use, with milk of lime, as Bordeaux mixture, to kill potato rot and grape fungus. Copper sulphate or blue vitriol, $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$, is a greenish blue crystalline substance. Other uses are in calico printing, plating, electrotyping, preserving timber, and purification of water.

343. Aluminum. The outstanding property of aluminum is its low specific gravity, or lightness. For this reason certain parts of airplanes are made of aluminum bronze, and auto bodies are frequently pure aluminum. In districts where labor and lumber are high, it is sometimes found desirable to use aluminum instead of copper to carry electricity. Its light weight makes it possible to get along with fewer supports. It conducts heat very well, and this combined with its luster and freedom from corrosion explains its use for cooking utensils. It is superior to copper, which is attacked by food acids, forming, sometimes, harmful compounds. Aluminum is not attacked by these weak acids. Bases, like sodium hydroxide, dissolve it. The equation for the reaction is



But we do not find strong bases in foods. Aluminum cooking utensils should not, however, be washed in strongly alkaline solutions, like lye and soda.

344. Zinc. The common name of zinc, on the market, is *spelter*. Its chief uses are in making brass, an alloy with copper, and in coating iron to make *galvanized iron*. On exposure to air, zinc becomes coated with a thin layer of basic carbonate, which adheres and so saves the rest of the metal. The container of a dry cell is made of zinc, which acts as one of the electrodes. A great deal of zinc is being used lately to make zinc white, an important

paint base. Like aluminum, zinc is amphoteric; that is, its compounds may act acid or basic. For example,



Here it acts acid, reacting with a base. But in the equation



it reacts with an acid, and therefore must be thought of as basic. Pure zinc is unattacked by most acids, but commercial zinc contains particles of carbon, which make with the zinc many small electrical cells; zinc loses electrons, becomes ionized and dissolves.

345. Tin. This is a silvery white metal, very seldom used alone, because of its low tenacity. Tin foil ("silver leaf") is nearly pure

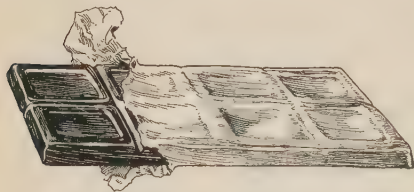


FIG. 231. Candy in a tin foil wrapper.

tin. Tin cans are made of iron plated with tin. This is true of nearly all so-called tinware. In fact, most of the tin used finds itself stuck to some kind of sheet iron. That is because tin is not easily oxidized or attacked by ordinary sub-

stances. For this reason, also, tin is used to line large copper utensils; and, although it wears off easily, it is also easily renewed by merely rubbing powdered tin on the hot clean copper surface. Tinned articles rust rapidly when once the iron becomes exposed; galvanized iron ware, therefore, is superior in wearing qualities, and should be used instead of tin, except where food is concerned. Tin forms many important alloys. Bronze is an alloy of tin and copper. Tin amalgam, an alloy containing mercury, is used to fill teeth. Solder is an alloy with lead. Pewter, britannia, babbitt, and the fusible alloys all contain tin.

346. Lead. Next to iron, lead is the cheapest metal, and if it were stronger, would be second in usefulness. Galena, PbS , is its chief ore. Its most important properties are its high specific

gravity, its softness and workability, and its resistance to corrosion. It is nearly always a constituent of fusible alloys, but its chief uses are in making pipe and type metal. Lead alloys and lead compounds are used in storage batteries. Most paint contains white lead, a basic carbonate, while red lead, an oxide, and chrome yellow, a chromate, are sometimes used as pigments.

347. Nickel. Pure nickel is very seldom used; its price is too high. It is used very largely to plate iron and brass articles, because it takes a high polish, and does not rust. Nickel alloys are very useful. German silver (Ni, Cu, and Zn), sometimes called nickel-silver, is practically a white brass. Silver-plated tableware is usually made of German silver with a silver coating. Nichrome, an alloy of nickel and chromium, has a high melting point and a high electrical resistance, which properties make it suitable for the heating elements of electrical heating devices — stoves, toasters, flatirons, etc. Nickel-steels are the most important iron alloys. They are harder and tougher than ordinary steels. They are used for steel tapes and other measuring instruments, as well as armored glass.

“Nickels” are made of 75 per cent copper and 25 per cent nickel. The metal may be made magnetic like iron, and in fact “permalloy,” composed of 80 per cent nickel and 20 per cent iron, is the most readily magnetized and demagnetized of all metals. It is therefore used in submarine cable work, and to make electromagnets.

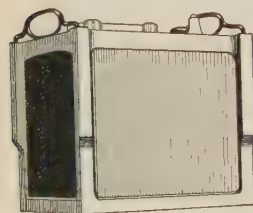


FIG. 233. A storage battery with lead plates.

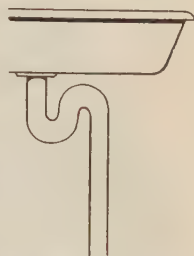


FIG. 232. Lead pipe.

348. Bismuth. The metal itself is brittle and has no uses. But it forms many useful alloys, of low melting point. Wood's metal, melting at 60.5° C., is the important part of the automatic fire sprinkler. Electrical fuses are made of bismuth alloys. Stereo-

type alloy, of such a low melting point that it can be cast in a paper mold, without setting it afire, is a bismuth alloy. Bismuth is



FIG. 234. Automatic fire sprinkler made of Wood's metal.

one of the few substances which expand on solidifying, so that its alloys make excellent clean-cut castings.

349. Antimony. This metal resembles bismuth rather closely. It is brittle, and therefore useless unless alloyed. But it has the peculiar property of expanding on solidification, and, when cast in a mold, it takes the exact shape of the mold. This is not true of most metals. Coins, for instance, cannot be cast, but

must be struck out by a die, because the metal, on cooling, draws away from the mold, and makes an imperfect impression. This property of antimony makes it valuable as a constituent of type

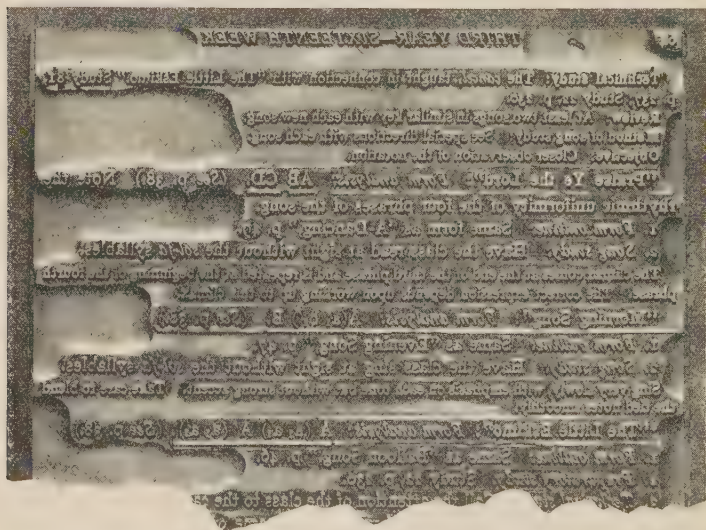


FIG. 235. Type metal.

metal. Other alloys of antimony are stereotype metal, britannia, and babbitt.

350. Magnesium. The properties of this metal which we find useful are its low specific gravity, 1.75, and its easy combustibility. When it burns, it produces a brilliant white light, which is useful in photography; it has about 10 per cent efficiency, — much higher than that of other luminants. Its light weight makes it a constituent of light alloys used in automobiles and airships. The metal is frequently used as a deoxidizing agent, in making alloys of copper and iron. It is of great value in organic synthesis, and is the only metallic constituent of chlorophyll.

351. Manganese. This element must not be confused with magnesium. The latter is silvery white, while manganese resembles iron. The pure metal is not used, but the alloys of manganese and iron, spiegeleisen and ferromanganese, are very important in making Bessemer and open hearth steels, and for certain hard, tough manganese steel alloys.

352. Mercury. Its most noticeable properties are its liquid state (mercury fuses at about -40° C., that is, it has a lower melting point than any ordinary metal), and its high specific gravity, 13.6. Its great weight, liquid state, and good conductivity make it useful for thermometers, barometers, air pumps, and other scientific instruments. It alloys with most metals, forming amalgams, which are exceedingly useful. An amalgamation process is used to recover gold and silver in the finely divided condition.

353. Silver. Silver is the best of the electrical conductors, but its price prohibits its use for this purpose, except in some delicate measuring instruments. Its chief use is in making tableware

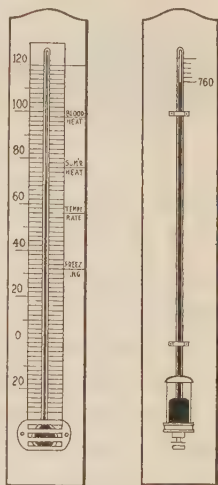


FIG. 236. A thermometer and a barometer. Both contain mercury.

(silverware), its exceptionally high luster being the desirable property. A great deal of so-called silverware has a white metal base with a silver plating. Silver tarnishes on exposure, due to the action of sulphur compounds (like H_2S in the air), which form Ag_2S , a black substance. This is frequently prevented by lacquering, that is, coating the metal with a thin layer of shellac. "Oxidized silver" is coated with Ag_2S by the action of Na_2S solution. Silver is very easily blackened by sulphur, as is shown when a rubber band is



FIG. 237. Silver coins.

kept in a pocket with a number of coins. American silver coins are 900 fine (that is, $\frac{900}{1000}$ silver); while British coins, sterling, are 925 fine.

Mirrors are sheets of glass, coated with pure silver, and then painted, to protect the metal. Some of the compounds of silver are of great importance. The entire business of photography depends upon the fact that silver compounds are sensitive to light. The halides of silver (AgCl , AgBr , and AgI) are the materials used. AgNO_3 , lunar caustic, is much used in medicine. When applied, it turns black, due to the formation of colloidal silver, and these stains are very difficult to remove. For this reason, it makes an excellent indelible ink, used in laundries.

354. Gold. Gold, the "king of metals," has some excellent properties; but its comparative rarity and high price forbid its common use. It is practically unaffected by natural reagents, like air and water; nor do ordinary acids attack it. Where small quantities are sufficient, therefore, as in dentistry, gold finds a ready and a useful market. It is too soft for ordinary use, however, and it is usually an alloy of gold that we use. Copper-gold alloys are reddish in color; silver-gold alloys are pale yellow, gray, or white; and alloys of copper, silver, and gold are green. Blue gold contains iron. The purity of gold is measured in carats; 24 carats being 100 per cent, so that 18 carat gold is $\frac{18}{24}$, or 75 per cent gold;

the remaining 25 per cent being, usually, copper. Nitric acid is used to test for gold. Brass turns green in nitric acid, due to the formation of $\text{Cu}(\text{NO}_3)_2$, while gold is unaffected by it. If the article is gold plated, it must first be scratched, to lay bare the metal underneath, before applying the acid.

Gold is dissolved by "aqua regia" (literally, kingly water; so called, because it dissolves gold, the king of metals), a mixture of nitric and hydrochloric acids, and forms gold chloride, AuCl_3 ; but when alloyed with silver, gold will not dissolve. When gold and silver are alloyed, they may be "parted" by a process called "inquartation." This consists in fusing the alloy, and adding sufficient silver to make the gold no more than one fourth of the total. After this, the silver may be dissolved out with HNO_3 .

Gold chloride reduced with stannous chloride gives a purple liquid, called "purple of Cassius," a colloidal solution of gold with stannic oxide as protective colloid.

American gold coins consist of 90 per cent gold and 10 per cent copper, while British coins contain 11 parts gold to 1 part copper. Gold is the most malleable substance we know. It can be hammered so thin that 1500 sheets of gold are about as thick as one sheet of writing paper. This gold leaf has many uses, such as edging books, sign lettering, designs on chinaware, etc. Brass articles are frequently plated with gold, for protection and for ornament; and where the plating is well done, it will last a lifetime. Such plated ware is usually superior in strength to similar articles made of "solid gold" of the same price. "Pinchbeck" is the name given to an alloy of 16 parts copper and 3 parts zinc, which, when new, resembles gold in color.

355. Platinum. The word "platinum" is derived from the Spanish adventurers of North America, who called it "platina," or "little silver," and considered it close to worthless. But in our time, the demand exceeds the supply, and the price has been driven up, until it is much more costly than gold. Its high melting point (1770°C.) and its resistance to attack by acids make it practically

the only substance out of which certain chemical apparatus can be made. Spark plug points and thermo-electric couples (for measuring high temperatures) are made of platinum or its alloys. Flame and bead tests in analytical chemistry are made with platinum wire; but of late the price has become prohibitive, and substitutes of one kind or another are being tried. Alloys of platinum or metals of the platinum family, particularly iridium and osmium, are exceedingly hard. Gold fountain pens are usually tipped with iridosmine.

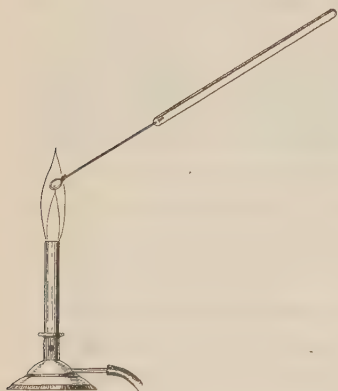


FIG. 238. Bead test with a platinum wire.

Platinum, and other members of its family, has the power to adsorb gases on its surface (occlusion). When a gas is compressed, heat is developed, which may be sufficient to start a flame if the gas is combustible, like hydrogen, illuminating gas, or gasoline vapor. This is the secret of the self-lighting burner. Of course, the greater the specific surface, the greater the adsorption;

we therefore expect fine platinum wire to act better than thick wire; but the best effect is produced by the use of colloidal platinum, precipitated in a silica gel. One volume of colloidal platinum adsorbs more than 100 volumes of oxygen, which is perhaps converted into the atomic variety, since it is more active than ordinary oxygen. This colloidal platinum is now being used in the "contact process" for the manufacture of sulphuric acid, to adsorb SO_2 and O_2 , and combine them into SO_3 . Platinum finds a similar use in the fixation of nitrogen; and it is probable that wherever gaseous reactions are to be brought about, the speed of the reaction can be increased very much by the use of an inert catalyst, like colloidal platinum.

Platinum alloys readily with the heavy metals, like lead and

bismuth; and as alloys of the heavy metals have relatively low melting points, they or their compounds should never be heated in platinum vessels.

356. Chromium is now being electroplated commercially on many different metal bases. It is practically tarnish-proof and imparts this property to articles protected by it. Its hardness is greater than that of steel and it is capable of increasing the wear of parts subject to abrasion. It resists nitric acid, salt water, sulphur compounds in petroleum, and a number of other corrosive agents.

Chromium is also used in making very tough alloy steels and stainless steel.

SUPPLEMENTARY READING

- E. E. Slosson, *Creative Chemistry* (Century Co., 1919), chap. xiv.
Svante Arrhenius, *Chemistry in Modern Life* (D. Van Nostrand Co., 1925), chap. iv.
H. E. Howe and F. M. Turner, *Chemistry in the Home* (Charles Scribner's Sons, 1927), chap. iii.
D. H. Killeffer, "Chromium Plating," *Industrial and Engineering Chemistry*, vol. XIX (1927), p. 773.

QUESTIONS

1. Explain why acid acts more rapidly on impure commercial zinc than on pure metal.
2. After answering question 1, explain the effect of amalgamating the zinc.
3. Mercury can be separated from its ore, HgS , by simple roasting. Why cannot this be done with CuS ?
4. What advantage has galvanized iron over sheet iron?
5. What effect has air and water on iron, zinc, copper, lead, and tin?
6. By example or otherwise, show that you understand the meaning of the words ore, gangue, smelting, roasting, reduction, mineral, noble metal, base metal, slag, flux.
7. Make a diagram of surface and underlying rock strata, showing where one would be likely to find ZnCO_3 and ZnS .
8. By what chemical process might silver be extracted from Ag_2S ?
9. What is a native ore? Name several native ores. Why are they found native?

10. How would an ore of composition PbO_2 be treated to get lead?
11. Make a diagram of the reverberatory furnace, and explain the extraction of lead.
12. Look up on the financial sheet of the daily newspaper the cost of five common metals. Convert the unit weight to pounds, and arrange them in order of increasing cost.
13. Make a tabular arrangement, showing five properties of metals, together with the name of the metal which best exemplifies each property.
14. What is meant by "pickling" metals? Why is it so called?
15. What methods are employed for protecting metals against corrosion?
16. Name in each case the compound formed when iron, copper, zinc, silver, and aluminum are acted upon by natural agents.
17. Select as many different metals as you can, and arrange them in the order of importance according to your individual experience. In a second column, give the names of articles made of each metal. (The importance of these articles should determine your arrangement.) In a third column, give the property of the metal which makes each article useful.
18. What is a "tin" pot; galvanized iron; German silver; "silver leaf"; white lead; gold leaf?
19. Why are storage batteries so heavy?
20. Tell how to distinguish a gold-plated from a solid gold article?
21. How can platinum be distinguished from silver?
22. What properties make platinum so valuable?
23. Name several countries that produce large amounts of iron; copper; gold; any other metal. What relation has this production to the prosperity of the nation?
24. Show how to recover gold and silver from a lead ore.
25. Mention metallurgical processes (methods for the extraction of metals) depending upon (a) reduction; (b) oxidation.
26. What is the composition of a "nickel"; a dime; a five dollar gold piece?
27. Why cannot any kind of wire be used in armored glass?
28. How much pure silver is present in a piece of sterling silver ware weighing 216 grams?
29. At market value, would it be worth mining a copper ore, containing 2 per cent copper, plus \$2.20 in gold, per ton, and \$1.15, in silver, per ton, if the cost of mining and extraction is \$5.60 per ton of ore?
30. What volume of SO_2 would be produced from 800 kilograms of ZnS , which is 70 per cent pure? What weight of H_2SO_4 would this make?
31. Referring to paragraph 44, select ten metals mentioned in this

chapter and, wherever possible, write equations between (a) metal and water; (b) metal and sulphuric acid.

32. Nitric acid is said to dissolve copper. Write the equation for this change, and show that the copper is really not dissolved as such.

33. Rewrite the reaction, $\text{HgS} + \text{O}_2 \longrightarrow \text{Hg} + \text{SO}_2$, in the electron form, pointing out which is the oxidizing and which the reducing agent.

34. Which metals are acted upon by alkalies. Write an equation in each case.

35. In recovering silver by replacement with copper, show by means of an electron equation that this is an oxidation-reduction reaction.

36. (a) How much zinc can be extracted from 10 tons of zinc sulphide, 75 per cent pure?

(b) How much lead can be extracted from 20 kilograms of lead sulphide, 100 per cent pure? What volume of air is needed for this operation?

37. How much gold is there in a 14 kt. ring weighing 15 grams?

CHAPTER XXVII

ALLOYS

357. A gold watch, an iron stove, a silver coin, a copper cent, — how often have we heard these expressions? And yet they are all incorrect, in the sense that the watch case is not made of pure gold, the stove is not pure iron, the cent contains tin as well as

copper. Very seldom, indeed, do we find any useful article made of pure metal, because the single metal has not the qualities desired. Nearly every metallic article is made of some alloy. What, then, is an alloy? To define an alloy is not so simple. It is easier to make it than to tell what it is. To make an alloy, it is, usually, only necessary to melt the two or more

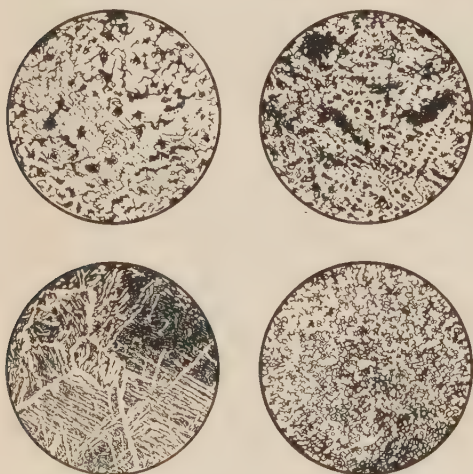


FIG. 239. Photomicrographs of alloys.

metals and mix them. While in the fused condition we have, of course, a solution. But what happens when the mass cools? As we learn in colloid chemistry, all liquids, which are crystallizing, pass through an intermediate colloidal condition, and this is the best stage in which to control the size of the resulting crystals.

In some cases the metals remain uniformly mixed, in which case we have an alloy which may be called a *solid solution*. In other

cases, the microscope shows that there has been a saturated solution, from which one or more of the metals have crystallized out. And then again, we have metals forming chemical compounds with one another.

The properties of alloys are dependent upon the size of particle or, in other words, upon the degree of dispersion. That is, crystalline structure corresponds to brittleness, while granular structure means elasticity and other valuable properties. The colloidal condition produces, usually, the best qualities.

358. Properties of alloys and constituents compared. When no chemical combination occurs on alloying, the properties are dependent upon the amounts of one or more of the constituents. But in most cases chemical combination does take place, and the alloy consists of a solid solution of this compound in one or more of the metals. Very little is known regarding these compounds, formulas for some of which are given here: Na_3Hg , NaHg_2 , Cu_2Zn_3 , CuZn_6 , Cu_2Zn_5 , Cu_3Al .

359. Hardness. Alloys are usually harder than their constituents; notable examples being gold and copper, silver and copper, copper and zinc, and the alloys of iron (steels) with chromium, nickel, and manganese. An alloy of $3 \text{ Pb} + 2 \text{ Bi}$ has ten times the hardness and twenty times the tenacity of lead.

360. Melting point. The melting point of an alloy is lower than that of one constituent and sometimes lower than the melting points of all the constituents. Brass has a melting point lower than copper, but higher than zinc. Some solders have a melting point lower than either lead or tin. Steel melts more easily than pure iron.

361. Specific gravity. The specific gravity of an alloy is not the average of that of its components. The following common

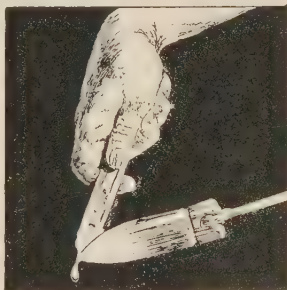


FIG. 240. Solder ($\frac{1}{2}$ lead, $\frac{1}{2}$ tin), whose melting point is lower than either lead or tin.

alloys have a specific gravity greater than might be expected: brass, bronze, type metal. Coin gold and coin silver have a specific gravity less than the average of the constituents.

362. Color. In some cases the color of the alloy is distinctly a surprise. Who would think that nickel coins are 75 per cent copper; that white gold contains 30 per cent gold? Green gold is an alloy of gold, silver, and cadmium. Blue gold contains iron. One can produce alloys of any desired color; the color being dependent, not on the constituents themselves, but on the degree of dispersion; that is, upon the size of the particle.

363. Chemical activity. Frequently, the alloy is much less active chemically. Sodium amalgam (NaHg) acts slowly on water, whereas sodium alone is quite vigorous, the amalgam acting more like a dilute solution. An alloy of 75 per cent gold and 25 per cent silver is not attacked by nitric acid; whereas, if there is at least 75 per cent silver in the alloy, nitric acid acts vigorously, dissolving the silver and leaving the gold.

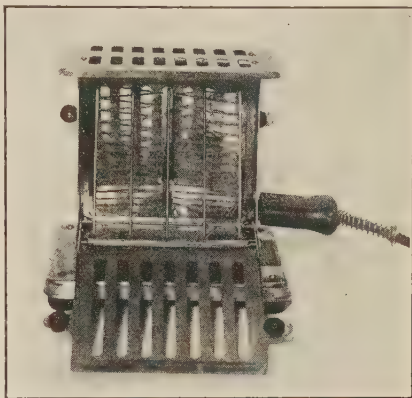
In commercial alloys, it is of course desirable to choose those with low chemical activity, so that corrosion is at a minimum. This is particularly so in the manufacture of iron articles. Stainless steels are alloys with high silicon or high chromium content. Some of these stainless steels can be boiled in nitric acid without appreciable corrosion.

364. Malleability of an alloy is usually less than that of its constituents; alloys tend to be brittle. For instance, pure iron (wrought) is the most malleable of all forms of iron, while some of the steels such as razor steel are exceedingly brittle.

365. Electrical conductivity. The electrical conductivity of alloys is always less than that of pure metals. Usually, very minute quantities of alloying material have a disastrous effect upon conductivity. One part of phosphorus in one thousand parts of copper makes the alloy useless for electrical purposes. German silver, an alloy of nickel, copper, and zinc, is used as resistance wire in rheostats. The "heating element" in electric stoves,

flat-irons, toasters, and similar devices is nichrome, an alloy of nickel and chromium. The alloy offers resistance to the current producing heat.

366. Annealing. When a metal or alloy is cooled rapidly, we get an exceedingly small grain, that is, a high degree of dispersion, which corresponds to increased hardness. Annealing, which means slow cooling, gives medium degrees of dispersion, where elasticity and toughness are at a maximum. For instance, when metals and alloys are rolled, they will tend to crack at the edges, unless first annealed; that is, heated to red heat and allowed to cool slowly. The process of tempering steels finishes with a partial annealing process.



Keystone View Co.

FIG. 241. An electric toaster with nichrome heating element.

367. Welding and soldering. Both of these processes produce alloys at the surfaces of contact, which must, therefore, be free from compounds. Fluxes and soldering pastes, used for this purpose, either reduce oxidation products to metal or remove them by fusion. Sandpaper or other abrasive may serve the same purpose.

SOME IMPORTANT ALLOYS

368. Amalgams. An amalgam is an alloy of some metal with mercury. In batteries the zinc is sometimes amalgamated to prevent "local action." The zinc amalgam acts as a cathode, while the carbon impurities of the zinc, unalloyed with mercury, are covered, and so prevented from forming electrical couples with the zinc. Tin amalgam was formerly used to make mirrors, and

it is still used to fill teeth. One part of tin rubbed together with four parts of mercury makes a plastic mass, which will set in a few days.

Gold is sometimes recovered from sands by amalgamation. The ground is washed away by water, which is then made to flow over copper plates, coated with mercury, to which the gold clings. This amalgam is then heated and the mercury distilled off.

Copper amalgam in boiling water becomes soft, while on cooling, it becomes so hard that it can be polished.

369. Fusible metals. All alloys have a lower melting point than one, and sometimes all, of their constituents; and advantage

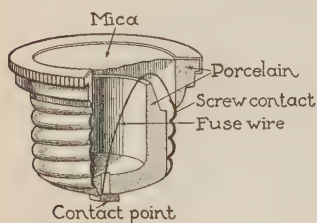


FIG. 242. A use for fusible alloys.

is taken of this principle to produce some alloys with very low melting points. Lipowitz metal (Pb, Sn, Cd, Bi) softens at 55°C . and is used to make casts. Wood's metal (another alloy of Pb, Sn, Cd, and Bi) melts below 60.5°C . and is used in automatic fire sprinklers (Fig. 234). The

water, shut off by means of a piece of Wood's metal, is released when the alloy is melted by hot gases arising from a fire. Automatic fire alarms work on the same principle. Safety plugs for steam boilers can be made to "blow off" at almost any desired steam pressure by using different alloys of lead, tin, and bismuth.

Electrical fuses have a strip of wire of fusible alloy, the low amperage fuse metal containing bismuth and cadmium, while the high amperage fuses have only lead and tin. When the current becomes great enough the heat melts the fuse and breaks the circuit. If there were no fuse, the copper wire within the walls would melt, and perhaps set fire to the house.

Type metal (Pb, Sn, and Sb) melts easily, so that it can readily be handled in a linotype machine. At the same time, the antimony (Sb) causes the alloy to expand on solidifying, thus insuring accurate type.

Solders are fusible alloys. Ordinary solders are alloys of lead and tin, "half-and-half" being a common one. Fine solder is two thirds tin and one third lead; it melts at 171°C . Gold solder contains gold, silver, and copper. Solders form alloys with the metals they join.

370. Bearing and antifriction metals. These are used to line bearings, to reduce the friction of a rotating steel shaft. At the same time, if the shaft should become overheated, then, instead of the steel losing shape, the bearing melts. This is preferable, since it is much easier to pour a new bearing than to replace a steel shaft. Babbitt metal, the most noted of bearing metals, is chiefly tin, with a little antimony and copper. This is frequently adulterated with lead, instead of tin, but the lead may easily be detected by its property of making a mark on paper.

371. In addition to the alloys which have already been mentioned, the following are also quite common: Magnalium (Mg and Al), used in airplanes because of its light weight; aluminum bronze (Al and Cu), used for jewelry and scientific instruments; duralumin (Al and Cu), used for household ware; brass (Cu + Zn), used for plumbing fixtures, door knobs, etc.; bronze (Cu + Zn + Sn), used for statuary, medals, and coins; lead shot, made of lead and arsenic.

A more extensive list of alloys will be found in the Appendix.

SUPPLEMENTARY READING

H. E. Howe, *Chemistry in the World's Work* (D. Van Nostrand Co., 1926), chap. vii.

QUESTIONS

1. Name some objects made of pure metal. Name some things made of alloys.
2. How is an alloy made?
3. Name some properties of an alloy which are an average of those of the constituents; some which differ from the properties of the constituents.

4. What is white gold; German silver; nichrome; solder?
5. What is meant by annealing?
6. Why is a flux needed before soldering?
7. What is an amalgam?
8. What is a fuse? Why should not a penny be used in place of a fuse?
9. Explain how a fire may be extinguished by an automatic fire sprinkler.
10. What important property is imparted to type metal by anti-mony?
11. Name the constituents of a bearing metal.
12. Choose from the table in the Appendix five alloys with which you have come into contact to-day, and arrange in a table the following information: name, composition, use, property upon which the use depends.
13. How can one distinguish silver from white gold and from platinum?
14. What would be the cost of material to produce one ton of brass (75% Cu + 25% Zn) if copper costs 14¢ per pound and zinc costs 6¢ per pound?
15. How much hydrogen can be obtained from one kilogram of sodium amalgam containing 10 per cent Na?
16. When rosin is used as a flux in soldering, the carbon and hydrogen act as reducing agents. Write oxidation-reduction equations to show what happens when Fe_2O_3 is reduced by (a) carbon; (b) hydrogen.
17. Name the products formed when nitric acid acts on (a) a silver coin (b) a copper coin; when aqua regia acts on a gold coin.

OPTIONAL QUESTIONS

18. Account for the enormous development of alloys in recent years.
19. What undesirable quality would be found in a brass which had been cooled quickly? How could that condition be remedied?
20. Examine your radio, enumerate all the alloys you can identify, give their composition, and state why, in your opinion, these alloys were used.
21. Compare the properties of an alloy with those of a solution.
22. Why does pure zinc not dissolve in sulphuric acid. Why does commercial zinc dissolve in acid? (Consider the electric cell.)

CHAPTER XXVIII

IRON

372. If we look about us and take stock of our material surroundings, we must notice that a very large number of articles are made of metal and that most of these are iron. In early history those men who developed iron weapons soon overcame their enemies who used only stone or bronze instruments. In Eastern Europe and Asia Minor, iron was apparently reduced 3000 years ago. In Spain, it was worked during the Roman occupation, and we are told that the early Britons opposed the Romans with iron spears and lances. Ever since, the march of civilization, as we know it, has kept pace with the development of the iron industry. Steel appeared rather early, but for centuries its making was shrouded in mystery, and the few men who could make it became famous. Even to-day, how many of us know the difference between iron and steel? What is a gas pipe made of? A stove? A knife? A nail? Why should one be made of iron and one of steel? What is the difference? To answer these questions intelligently we must study the subject and the time spent will be amply repaid. How often do we find articles made of cast iron which should be steel or wrought iron! The choice of the proper material frequently makes the difference between success and failure in many an enterprise. We must know. We cannot afford to guess.

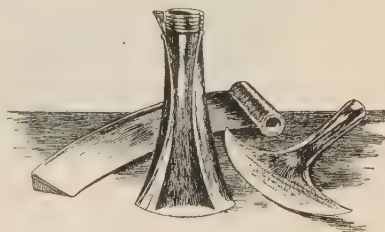
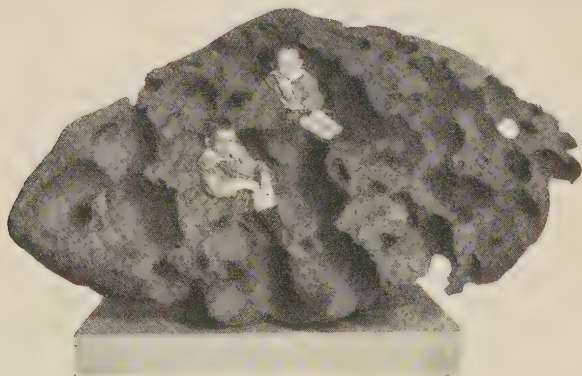


FIG. 243. Early iron implements.

373. Sources of iron. Iron is one of the most widely distributed of the elements. A little free iron is found in particles in rocks and in meteorites, but its numerous compounds are scattered throughout the world of plants, animals, and rocks. There is more iron in the world than any other metal except aluminum. One twentieth of the earth is composed of iron. In animals it exists in the form of hemoglobin, in the red blood cells. It is present in plants, and it has been shown that the oxidation processes carried on in the body are impossible without the presence of iron, and



Museum of Natural History

FIG. 244. Meteorite made of iron.

hardly a mineral exists which does not contain some iron. Thus to mention a few : pyrites, or fool's gold, asbestos, mica, hornblende, magnetite (lodestone), hematite, limonite, and siderite all contain iron. The brown, red, and yellow tints in rocks and soils are caused by iron compounds, as is also the brown deposit in water.

374. Ores of iron. While iron is so common, it is not always easy, or economical, to extract it from some of these materials. Therefore, since the extraction, or *metallurgy*, of iron is in the hands of business men, who must make a profit, they use only those substances which are found in large quantity, which can be easily mined, and which can be easily transported and treated. These are called *ores*. The ores of iron are :

MINERAL NAME	OTHER NAMES	CHEMICAL COMPOSITION
hematite	red hematite; specular iron ore	Fe_2O_3
limonite	brown hematite; bog iron ore	$2 \text{Fe}_2\text{O}_3 \cdot 3 \text{H}_2\text{O}$
magnetite	lodestone	Fe_3O_4
siderite	spathic iron ore; clay ironstone	FeCO_3

Magnetite is found in large quantity in Sweden, where the ore is so pure that it is used by the famous steel makers of England, although there is abundance of iron mined in that country. In the United States we mine magnetite in New York and New Jersey. Siderite is the chief ore of England, where it is known as clay ironstone, but it is of no importance in the United States.

The chief ore of iron is hematite. It is the principal source of the iron of Germany, France, and Belgium, and of our own enormous deposits in Minnesota, Michigan, and Alabama.

375. Handling ore. In the Lake Superior region, our hematite is mined by steam shovel where the deposit is near the surface; and frequently by "caving" where it is underground, the ore being run out of hoppers into cars on underground tracks. It is drawn to the surface, dumped directly into freight cars, which are then drawn to the lake ports, Duluth, Marquette, etc., where the facilities for handling are excellent. Docks, a thousand feet long, with railroad tracks on top and enormous bins underneath, tower above the ore vessels drawn alongside. Each dock can load four vessels at a time by means of long chutes, let down from the bins in the dock, into the vessels. The ore is carried to Detroit, Cleveland, Erie, Buffalo, and other towns along the Great Lakes, where it is unloaded into freight cars, which carry it to the furnaces, situated about Pittsburgh, New York, Buffalo, and other centers of the steel industry. Here again the ore is dropped into bins, at

the furnaces, and hoisted by "skips," little cars, to the top of the furnace. The entire manipulation is mechanical, the watchword being economy. The steel industry has centered about Pittsburgh because the coke, necessary for the reduction of iron,

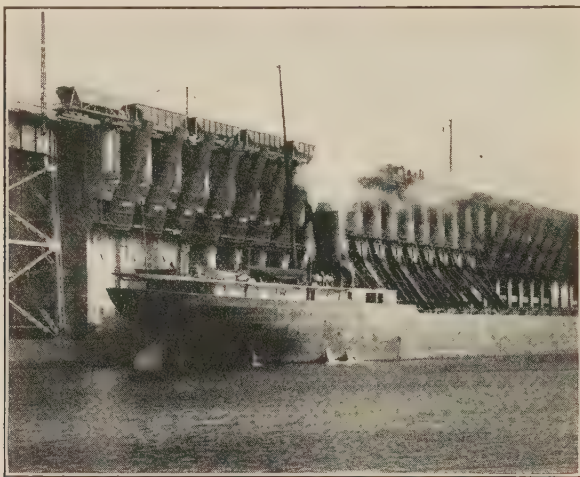


FIG. 245. Vessel being loaded with iron ore at a lake port.

and which is produced in Pennsylvania, will not bear transportation, as it powders easily. Therefore the ore must be carried to the coke.

376. Manufacturing of cast iron. It will readily be understood that material mined with a steam shovel, or by "caving," cannot be pure. Pieces of rock and earth will be taken with the ore with which it is sometimes intimately associated. This foreign matter is called the "gangue," and consists, usually, of sand, clay, and rock. The gangue must be eliminated, and for that purpose there is added a calculated weight of limestone, CaCO_3 , which unites with the silica to make a fusible mixture of silicates, like glass. This product is called *slag* or cinder. The calcium carbonate makes the solids "fluid," so it is called the *flux*. The ore is an oxide, Fe_2O_3 ; therefore something must be used to remove

the oxygen, — some reducing agent. Coke is found best; it is cheap, fairly pure, and is porous, allowing air to get through.

377. Blast-furnace operations. The ancients smelted iron in the open, using a wood fire; but their processes were so inefficient that, for example, the slags discarded by the workers of Dent forest, in England, have furnished ore for twenty modern furnaces for 250 years.

Modern smelting of iron starts in a blast furnace, the product of which is called *pig* or *cast iron*. The modern blast furnace is cylindrical in shape and very large. Its dimensions may be 30 feet in diameter and 100 feet high. The outside is a shell of riveted wrought-iron plates, lined with a thick layer of fire brick. The hearth, or crucible, constituting about ten feet of the lowest part of the furnace, is lined with an exceptionally thick layer of brick, in which pipes are set, so that it may be water cooled. Right

above the hearth are the *tuyères* (pronounced “tweers”). These are pipes (protected by firebrick, water cooled) through which a blast of preheated air is forced under high pressure. The furnace is run 24 hours a day, 365 days per year, until the lining is worn so badly that repairs are necessary. It may last two to twenty years.

The ore, the coke, and the limestone are analyzed, and the proportions necessary for the charge are calculated. The charge consists, roughly, of:

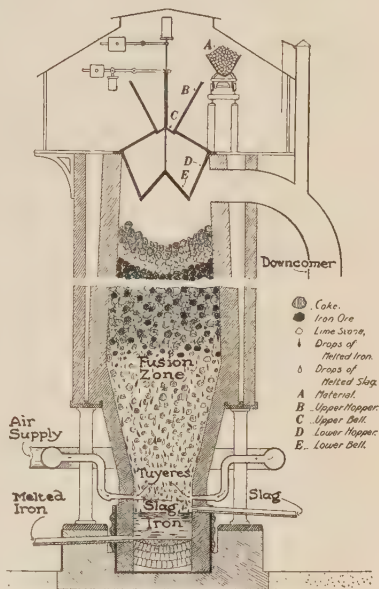
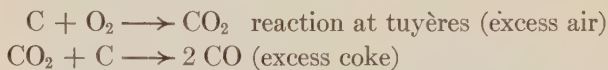


FIG. 246. Blast furnace.

limestone	1 part
coke	2 parts
ore	4 parts

These are dumped into the furnace by means of a double "cup and cone," shown in the diagram. This device prevents the escape of gases during the operation of the furnace. The coke burns at the tuyères, producing carbon dioxide, which is reduced to carbon monoxide by the excess of coke according to these equations:



It is this carbon monoxide which accomplishes the chief reduction of the Fe_2O_3 :



Carbon also reduces Fe_2O_3 , so that we may write



The molten iron gradually sinks, absorbing carbon, sulphides, and other things on its way, and forms a layer beneath the charge. While this is going on, the limestone combines with the silica of the gangue and the ash of the coke to form a silicate:



and this calcium silicate fuses with the aluminum, iron, and magnesium silicates of the gangue to form slag, which lies on top of the molten iron in the crucible.

Slag is tapped every four hours to prevent its rising to the tuyères, and the iron is tapped about four times a day, a hundred to a hundred and fifty tons at a time. Some furnaces produce five hundred tons a day. This molten iron is run into a bed of sand molds, each about two feet long, and when cool, these are broken apart by sledges, into "pigs." The gases issuing from the furnace contain:

N_2		70%
CO		20%
CO_2		10%

This is similar to producer gas (par. 283), and it is used to heat the hot-blast stoves, — enormous structures, four or five to a blast furnace, filled with a checkerwork of brick or flues. The gases may also be used to furnish fuel for steam and gas engines to run blowers, pumps, and elevators. Some slags are used to make cement, glass, and sometimes fertilizer; but the utilization of the enormous quantities produced is still an unsolved problem.

When the ore is rich in manganese, the cast iron made is called “spiegel” if it contains 20 per cent Mn; and ferro-manganese, if it contains 80 per cent Mn. These are used in the manufacture of steel.

The following tabular outline will perhaps assist in understanding the manufacture of cast iron:

MATERIAL USED	COMPOSITION	REASON FOR USE	MATERIAL FORMED
Ore	Fe_2O_3	contains iron	Fe
Gangue	SiO_2	mixed with ore	slag — $CaSiO_3$, etc.
Coke	$\left\{ \begin{array}{l} C \\ SiO_2 \\ FeS_2 \end{array} \right.$	fuel and reducing impurity	$CO + CO_2$ slag
Limestone	$CaCO_3$	flux	FeS in iron slag

VARIETIES OF CAST IRON	COMPOSITION	PROPERTIES
White	Carbon combined with iron as Fe_3C	Very brittle but hard
Gray	Carbon free as graphite	Not so brittle and softer than white
Malleable castings	Cast iron inside, carbon removed from outer layers	Less brittle, and tougher than either

378. Composition of pig iron. A characteristic analysis of pig iron is here given (in approximate figures) :

Fe	93.0%
C	4.0
Si	1.5
S	0.5
P	0.5
Mn	0.5
	100.0

It should be noted that the impurities (everything besides iron) are rather high — about 7 per cent.

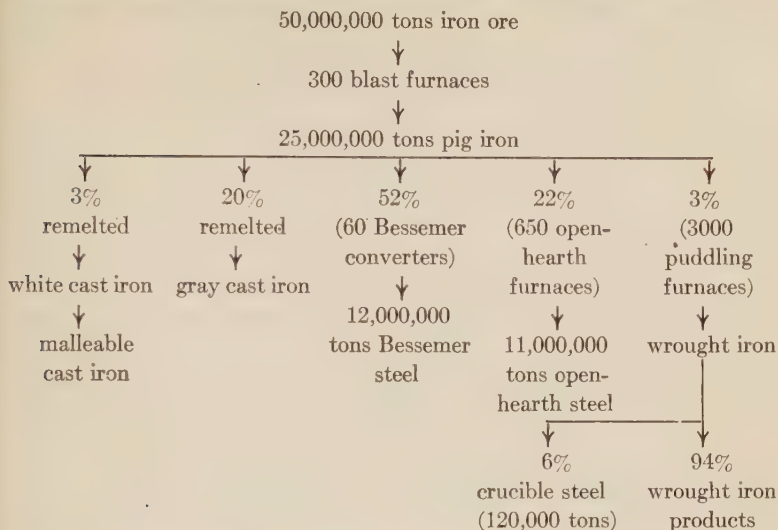
379.

PROPERTIES OF CAST IRON	USES DEPENDENT ON PROPERTIES
Expands slightly on solidifying	castings { ranges pipes radiators bathtubs ornamental objects
Resists pressure (not bending or twisting)	
Heavy and cheap	
	columns
	engine foundations
	steam rollers

Besides these properties which make cast iron very useful, there are some which make it worthless for many purposes. It is exceedingly brittle, breaking easily under stress of bending, torsion, and tension, while the blow of a hammer sometimes breaks it almost as easily as glass.

380. Definition of cast iron. Cast iron may be defined as iron containing so much carbon that it is not malleable at any temperature. It is evident that where great strength is required against all kinds of stresses, some other material must be used. For these reasons, most of the pig iron produced is converted into

other alloys of iron, as will be shown by the following table adapted from Stoughton's "Metallurgy of Iron and Steel."



381. Manufacture of wrought iron. To render cast iron less brittle, the impurities, chiefly the carbon, sulphur, and phosphorus, must be removed. This is accomplished by oxidation and slagging in a special type of reverberatory furnace, called a puddling furnace, because the molten iron is worked with iron rods: "puddled."

The chief differences between the reverberatory and blast furnaces are two: the

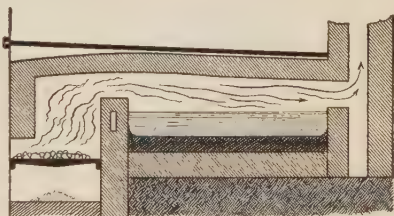


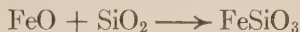
FIG. 247. A reverberatory furnace.

reverberatory has an oxidizing atmosphere, while the blast furnace has a reducing atmosphere; and whereas the fuel in a blast furnace is mixed with the ore and iron, while in the reverberatory, the fuel is in a separate firebox, reducing contamination to a minimum.

382. Details of the process. The pig iron is melted on the hearth lined with pure hematite, Fe_2O_3 . In about a half hour the molten mass "boils," due to elimination of carbon monoxide, which burns with a blue flame over the iron.



The silicon is slagged by FeO to form iron silicate:



The sulphur and phosphorus are oxidized and eliminated chiefly as sulphates and phosphates by the basic lining of the hearth, Fe_2O_3 . In one and one half hours the sulphur and phosphorus have been eliminated, and the carbon is reduced from 4 per cent to 0.1 per cent. At this stage the mass becomes pasty, because the purer iron has a higher melting point than the impure iron (1200°C . for cast iron, 1600°C . for wrought iron). It is then puddled or worked with iron bars into several balls, called "blooms," weighing about 100 lbs. apiece. This iron is quite pure except for the slag which is retained by the pasty iron. In order to remove as much of this as possible, the bloom is squeezed under huge trip hammers, the slag oozing out. It is then rolled into a bar, cut into strips, welded, and again rolled to remove more slag and make a network of slag fibers.

Wrought-iron pipe is made by bending the rolled strip into shape, then heating and welding by drawing it through rolls. Wire is made by drawing a hot wrought-iron bar through a die (piece of steel with a hole) each successive die being smaller. The composition of wrought iron is (approximate figures):

Fe	99.6%
C	0.1
Si	0.05
S	0.05
P	0.15
Mn	0.05
	100.00

Note the low percentage of impurities — less than one per cent.

PROPERTIES	USES
malleable and tough	sheet iron for roofing, etc.; horseshoes and other hammered ware; "hardware," such as nails, bolts, hinges, locks, pipe for threading; chains; anchors; agricultural implements
ductile	wire; I bars for bridges
soft	keys and small pipe
magnetic	cores of electromagnets for motors, dynamos, telephone receivers, etc.

383. Definition of wrought iron. Wrought iron is a slag-bearing malleable iron, which does not harden materially when suddenly cooled. Since wrought iron is expensive, low-carbon steels, which have nearly all the desirable qualities of wrought iron, have been substituted.

384. Steel. Being intermediate in composition between cast and wrought iron, steel may be made from either — from wrought iron, by addition of carbon; and from cast iron, by removal of part of the carbon as well as impurities. The ancients made steel from

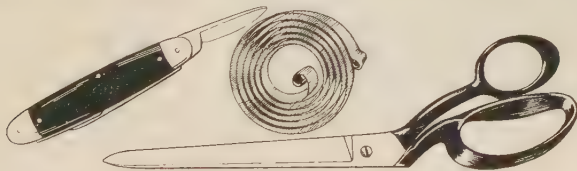


FIG. 248. Crucible steel articles.

wrought iron; and, until the nineteenth century, this was the chief process. The method still in use among us, which corresponds most closely to the ancient process, is the *cementation* or *case-hardening* process. Bars of wrought iron are packed in earthenware boxes with charcoal and heated, 1000° – 1100° , for ten days. The iron absorbs about 1 per cent carbon. The *crucible process* produces steel resembling this in composition. Cast iron and wrought iron are fused together in a graphite crucible. To give

the proper percentage of carbon and to deoxidize the steel ferro-manganese is added. This is sometimes called *cast steel*. Personal attention and care can be given to this steel and it is therefore well made and free from gases; but only small articles can be made of it, since it is very expensive. Cutting tools, like knives and razors, springs, needles, pens, and similar articles are made of crucible steel.

385. Bessemer steel. Since large amounts of steel cannot be handled in a crucible, three processes have been developed for making large articles. One of these is known as the Bessemer process. The idea underlying this method is first to burn out all the impurities from cast iron and then "recarburize" (add carbon again) to the desired extent. Air is used to oxidize carbon, sulphur, silicon, and phosphorus compounds, some of which pass off as gas, while

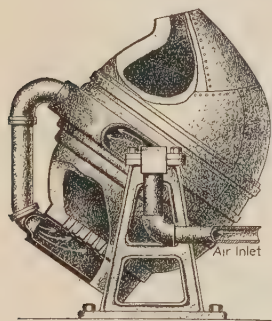


FIG. 249. A Bessemer converter.

the rest are slagged. This leaves pure iron, to which is then added spiegel or ferromanganese of known composition, so as to produce a steel with the desired percentages of carbon and manganese. The manganese, being more active than iron in combining with oxygen, prevents blowholes caused by the escape of gases. The manganese also has an important effect in making the steel self-tempering, as we shall see later under that head.

386. Details of the process. The Bessemer converter is like the blast furnace, in that a terrific blast of air is blown through the charge; but no fuel is used. The oxidation of the impurities of the cast iron causes sufficient heat to keep the charge molten. The converter is unlike the blast furnace in that it is small, holding only fifteen tons of melted pig iron, and it is supported on shafts so that it can be rotated for loading and emptying. It is, like the blast furnace, a pear-shaped, iron vessel, lined with firebrick, containing sometimes ground limestone and dolomite (CaCO_3 plus

MgCO_3) to assist in removing sulphur and phosphorus compounds. This is the *basic lining* method devised by Thomas and Gilchrist.

The molten pig iron, direct from the blast furnace, is poured into the converter while on its side, the air is turned on, and the converter is raised, mouth up. Thirty thousand cubic feet of cold air per minute are blown in, and in four minutes all the silicon and manganese are oxidized.

Then carbon burns to carbon monoxide, which flames up. In six minutes the flame drops, showing that the carbon has been oxidized. The temperature is now higher than at first, as it must be, to keep the material molten, since it is now like wrought iron: pure but infusible. Just as soon as the carbon is burned out, the converter is turned on its side and the air shut off; otherwise the iron would be blown full of holes and oxidized.

The recarburizing is now done, by adding the calculated weight of *spiegel*; the mass is poured at once into ingots and rolled while hot. The process is rapid, taking about ten minutes per "heat." One converter can handle the product of two blast furnaces.

387. Composition of Bessemer steel. It contains from 0.6 per cent carbon, for soft steel, to 2.0 per cent carbon, for hard steel; together with 1.5 per cent manganese; while silicon, sulphur, and phosphorus together are less than 0.1 per cent. Even small percentages of sulphur or phosphorus have very bad effects on the



Underwood and Underwood

FIG. 250. Bessemer converter in action.

steel, making it brittle. The low carbon Bessemer is practically wrought iron, and is replacing wrought iron for many purposes. Bessemer steel is used for rails and structural iron, such as is used for buildings and bridges; but the product is not very uniform in quality, frequently containing blowholes; and furthermore, the so-called Bessemer ores, low in phosphorus and sulphur, are getting scarce and expensive, so that this process is giving place slowly, but surely, to the open-hearth process.

388. Open-hearth process. This resembles the Bessemer process in general idea; the impurities are oxidized and driven off, and the carbon and manganese added in desired quantity. But the open-hearth process produces a better and more uniform product.

389. Details of the open-hearth process. From fifteen to one hundred tons of pig iron, or scrap iron, or both are melted in a reverberatory furnace, heated by gas; but *no air is blown through the charge*. Oxidation is accomplished by adding iron ore, Fe_2O_3 ; and when there is much sulphur and phosphorus, these are slagged by using a basic lining. In about two hours the charge is melted and samples are sent up for analysis. Six to ten hours constitute a "run," during which the process is carefully watched and regulated by constant analyses. After the carbon, sulphur, silicon, and phosphorus are burned out as far as desired, the charge is recarburized with spiegeleisen, the same as in the Bessemer process. The quiet heating, with absence of air blasts, gives freedom from blowholes, and on the whole this steel is a superior product. It is used for better class rails, shafts, armor plate, guns, and structural steel.

390. Electrical treatment of iron and steel. There are processes for the electrical treatment of ores to produce pig iron, using charcoal as a reducing agent, and they give an excellent product; but as yet they are too expensive to compete with the older methods. But as a method of refining, the electric furnace has come to stay.

It usually takes metal from the Bessemer or open-hearth processes with impurities burned out, before recarburization. Fe_2O_3 is added to oxidize impurities still further and to slag the phosphorus; if necessary, the slag is skimmed and more flux added. Sulphur is removed by lime in the presence of carbon. Gases are removed, and therefore blowholes prevented, because no air blast is used, and the process is quiet, allowing the gases to rise slowly and escape. No fuel is used; impurities are therefore left out. But this steel is more expensive than open-hearth steel. Still the quality of the product is as good as, and sometimes better than, crucible steel, and much cheaper. Without doubt, the near future will see crucible steel replaced by open-hearth steel refined electrically.

391. Definition of steel. Steel may be defined as iron which is malleable at least at one temperature, and which can be hardened greatly by sudden cooling.

392. Composition of steel. A typical analysis of steel shows (approximate figures):

Fe	98.98%
C	0.50
Si	0.10
S	0.01
P	0.01
Mn	0.40
	100.00

Note the low sulphur and phosphorus and the medium carbon.

393. Properties of steel. Steel has a range of properties which is unequalled by any other substance. It can be made hard or soft, malleable or brittle. It can be made like cast iron and wrought iron, and it can assume the properties of many other substances. It can be permanently magnetized and it has, in the highest degree, the valuable property called elasticity. These properties of steel make possible our present-day civilization, with its railways, steamships, bridges, etc.

394. Comparison of cast iron, wrought iron, and steel.

%	CAST IRON	STEEL	WROUGHT IRON
Fe	93.0	98.83	99.60
C	4.0	0.50	0.10
Si	1.5	0.10	0.05
S	0.5	0.01	0.05
P	0.5	0.01	0.15
Mn	0.5	0.55	0.05
	100.00	100.00	100.00

A study of the typical analysis reveals the fact that steel is intermediate in composition between cast iron and wrought iron, in

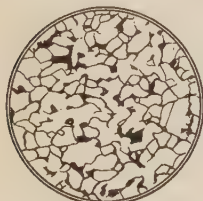


FIG. 251. Cast iron under the microscope.

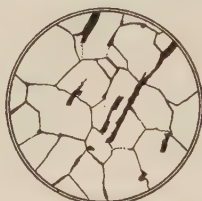


FIG. 252. Wrought iron under the microscope.

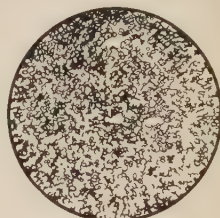


FIG. 253. Steel under the microscope.

respect to carbon, the chief impurity; lower than either in sulphur and phosphorus, higher in manganese. It is, however, impossible to detect a steel by its analysis alone. Some

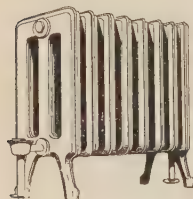


FIG. 254. Cast iron.

low-carbon steels are practically wrought iron, while some are no better than the electrical cast irons. *Malleable iron* should be wrought iron; but very frequently it is cast iron, toughened, on the surface, by removal

of carbon. Under the microscope, wrought iron shows a characteristic fibrous structure, due to the thin lines of slag drawn out during rolling. Cast iron will show relatively large crystals while steel will show small crystals. Steel is a true colloid.



FIG. 255. Wrought iron.

The ingredients present in cast iron, wrought iron, and carbon steels and their effects are :

INGREDIENT	NAME	PROPERTY GIVEN TO IRON
Iron	ferrite	malleable, soft
Fe_3C	cementite	hard, brittle, scratches glass
FeS	sulphur	brittle
Fe_3P	phosphorus	brittle
FeSi	silicon	increases strength and hardness ; reduces oxygen
MnS	manganese	makes smaller crystals, but makes metal brittle when hot
CO , FeO , Fe_2O_3	oxides	brittle
C	carbon-graphite	brittle

395. Tempering. The relative amounts of carbon and iron present as graphite, cementite, and ferrite, and their degree of dispersion, or colloidal condition, determine many of the properties of iron and steel ; and changes in these conditions are brought about by heat treatment and also by the presence of certain elements. *Temper* means hardness, produced chiefly by cementite, Fe_3C . However, if the crystals of Fe_3C are large, brittleness results. Add to this the fact that below 690° , the *critical temperature*, the solid solution of cementite, ferrite, and carbon begins to crystallize, passing through the intermediate or colloidal condition, and we have all that is necessary to understand the mysterious tempering process (so often associated with "magic art" in the middle ages. Frequently, a smith, who had extraordinary skill in tempering steel, would attribute his success to supernatural means: performing the task under a full moon, reciting a certain formula, etc.)

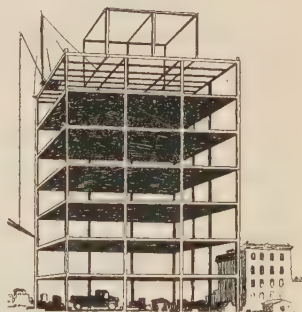


FIG. 256. Steel.

If iron and carbon are heated together to white heat, a molecular solution of cementite in ferrite results, and if this is cooled quickly (quenching), it remains molecular; and this kind of steel has maximum temper. If cooled slowly, as soon as it passes 690° C. separation of ferrite takes place and we get soft iron, large cementite crystals, and flaky graphite: which means minimum temper. Most tempered steels are somewhere in between these two extremes. How is that accomplished? We must have cementite, but it must be in colloidal condition, and not at either of the extremes; that is, it must not be molecularly dispersed or coarsely crystalline. Neither must the carbon exist in large crystals, *i.e.*, flaky. This may be brought about in one of two ways: (1) by cooling slightly below 690° before quenching, or (2) by quenching at 690° and then reheating at lower temperatures. In either case more time is allowed, below 690° , for the molecular mass to coagulate into larger aggregates, so that we shall have the solid solution called steel. At any desired moment, further change can be prevented by sudden cooling, because at ordinary temperatures, practically no changes are taking place. We may now, perhaps, give a new definition of steel, as consisting of cementite in colloidal solution in ferrite. Modern practice in tempering requires reheating at lower temperatures, 225° to 300° C., because these temperatures correspond to the formation of oxides of iron in thin films, on the surface, and a play of colors is produced (same reason as for oil on water or soap bubbles) corresponding to the thickness of this film. An experienced man can tell, then, without a thermometer or other instrument, when the reheating is finished, because each color corresponds to a certain temper. However, time also has an important effect in the formation of crystals and this is usually overlooked. In the table may be seen the temperature corresponding to various colors in the reheating, together with articles made at each temperature. It will be noted that the steel gets softer, the longer it is reheated: at 225° , drills (hard) are tempered; while at 285° , we get wood saws (soft steel).

COLOR	TEMPERATURE	ARTICLES
very pale yellow	225° C.-235° C.	steel engraving tools hammer faces drills
straw yellow	235° C.-250° C.	rock drills penknives planes dies shears stone-working tools
yellow brown	250° C.-265° C.	wood-working tools
light purple	265° C.-275° C.	augers dental and surgical instruments
dark purple	275° C.-285° C.	axes gimlets needles hack saws screw drivers
blue	285° C.-300° C.	springs saws for wood

Cast iron is subject to tempering, although the term is not usually applied in that connection. White cast iron has a hard temper, due to sudden chilling, while the gray variety is softer and less brittle. The cooling of castings must be regulated, therefore, to insure the proper temper, or they will be too soft or too hard, and certainly not uniform. While small articles may be reheated to temper them, that becomes impossible with large units like structural steel. Such articles were therefore impossible of manufacture, until the discovery of manganese steel (Bessemer), which is self-tempering. This is due to the effect which manganese has on the critical temperature (690° C.), lowering it below ordinary temperatures; so that the decomposition of the solid solution of cementite and ferrite does not occur, as it does in ordinary carbon

steel, below 690° , and the steel does not lose its temper (get soft) by slow cooling.

Tungsten and molybdenum also have this effect on the critical temperature, and in fact they widen the range of critical changes so much that the steel may be heated to redness without losing temper.



FIG. 257. Rail curve made of nickel-steel.

396. Alloy steels. Steel may be made to assume almost any physical property by alloying with other metals, and some of these alloy steels are of very great commercial importance.

397. Nickel-steels. Most important of all are the nickel-steels. The addition of 3.5 per cent nickel increases the elastic limit of steel 50 per cent and it remains elastic six times as long. It is used for rail curves, ordnance, bridge steel, shafting, parts of automobiles and rail cars, axles, etc. It is both harder and tougher than ordinary carbon steel, has less friction, and no blowholes.

Invar, containing 36 per cent nickel, does not expand or contract when heated or cooled, and is used for pendulums, steel tapes, and other scientific instruments.

Platinite, with 45 per cent nickel, has the same coefficient of expansion as platinum and glass. It is used for armored glass (nickel-steel wire fused into glass).

398. Manganese steel. Manganese up to 12 per cent is added to steel to make a hard but not brittle alloy used for rock crushers, railroad switches, curves, and safes. It is of limited use, because it is so hard that it cannot be worked, but must be cast into its final shape. Ordinary carbon steel can be tempered



FIG. 258. Pendulum made of invar, a nickel-steel.

and annealed so it is shaped before tempering. Manganese steels, however, do not lose temper, and so are hard from the time of casting. They cannot be annealed. One disadvantage of manganese steels is that their elastic limit is very low, making them poor for structural work.

399. Chrome steel. One to two per cent of chromium added to a carbon steel makes a very hard alloy with high elastic limit. It is used for armor plate, crushing machinery, and automobile gears. Fifteen per cent chromium produces stainless and rustproof steel and iron.

400. Silicon steel. This usually contains 1 to 5 per cent silicon, and the smallest amounts of carbon, manganese, etc. It has better magnetic properties even than pure iron and is much used

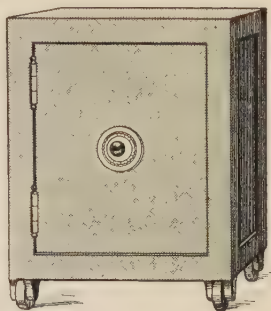


FIG. 259. Safe made of manganese steel.

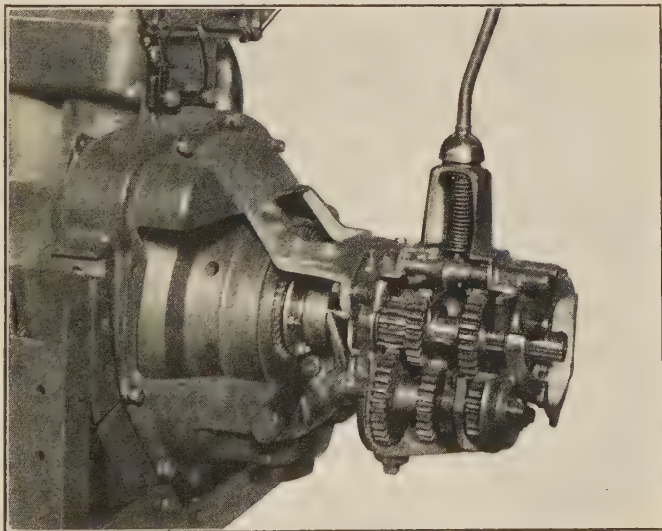


FIG. 260. Automobile gears made of chrome steel.

in electrical work. The 13 per cent silicon steel resists rusting by acids and is used for *stainless steel*.

401. Vanadium steel. One tenth to one fifth of one per cent of vanadium has marked effects in deoxidizing steel and preventing blowholes. It increases the strength and duration of elasticity.



FIG. 261. Automobile spring made of vanadium steel.

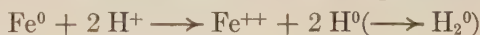
Vanadium steels are used for axles, connecting rods, gears, solid wheels for railroad cars, and automobile springs.

402. Tungsten steel. Low-carbon steels are alloyed with tungsten up to 24 per cent, to increase hardness and toughness at the same time, and make self-hardening alloys which do not lose their temper even at red heat. Such alloys are used to make tools for high-speed lathes where metals are worked. The tools get red hot, sometimes, and if made of ordinary carbon steel would lose their temper, while tungsten steel remains hard.

403. Molybdenum steel. Half the quantity of molybdenum seems to produce the same results as tungsten.

404. Rusting of iron. Millions of dollars are expended yearly to fight the attacks of natural agents on iron, and thousands of tons of iron and steel are annually turned to rust on the insides of pipes, boilers, and tanks, where the iron is not protected. The problem of protecting iron must be attacked with full knowledge of what are the agents which corrode it, how they act, and what they produce.

Rust is of varying composition. It starts as a hydrated (containing water) mixture of ferrous and ferric oxides, with ferrous carbonate, gradually changing to Fe_2O_3 . In general, acids cause iron to rust, while alkalis retard the corrosion. The action of rust formation seems to be due to the fact that iron, being above hydrogen in the electromotive series, tends to lose electrons and enter solution; so that in water or acid, we get



Oxygen removes this hydrogen as fast as formed, thus upsetting the equilibrium, which therefore proceeds to the right. When this reaction is complete (when the iron or the water is used up), the ferrous hydroxide, $\text{Fe}(\text{OH})_2$, is oxidized to ferric hydroxide, $\text{Fe}(\text{OH})_3$, both at the same time combining with carbonic acid to form ferrous carbonate, and these are slowly decomposed to form ferric oxide, Fe_2O_3 .

Rust stains may be removed by washing with dilute vinegar or lemon juice, or better, oxalic or hydrochloric acids. These change the iron rust to soluble compounds, like FeCl_3 , which must then be washed away by water; otherwise on exposure to air the stain will return.

405. Protection of iron. There are many ways to prevent the corrosion of iron, and these usually consist in keeping iron and its enemies apart. In every case the iron should first be cleaned, otherwise the rust, which is exceedingly brittle, will crumble, carrying the protecting layer with it. Before painting, iron is prepared by sand blast or steel brush. Before plating, it is "pickled" (washed in sulphuric acid), which dissolves rust to form soluble ferrous sulphate.

Where paints are used, there is no particular choice, except that none of the ingredients of the paint should be of acid character, or in any way attack iron.

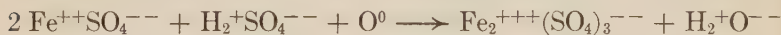
Zinc, tin, and nickel are the chief metallic protective agents used for iron. Nickel is by far the best of these, but its cost forbids its use in most cases. As between zinc and tin, the zinc is far superior, both as regards cost and protection, and should always be selected except where articles of food are concerned, and then the tin is to be preferred. If tinware once exposes a spot of iron, the corrosion proceeds rapidly; whereas, with zinc, the protection lasts while there is a bit of zinc left. This is readily understood by reference to the electrochemical series of elements. Zinc is positive to iron, and as we have seen, when both are present in a medium with which they can react, the zinc is attacked, and the

iron is unaffected. This must continue until all the zinc is gone, when the iron will be attacked in its turn. Tin, on the contrary, is negative to iron, so that when tin and iron are present the iron is attacked first. Some engineers hang pieces of zinc in their boilers to protect the iron.

406. Other ways to protect iron. *Enamel.* A coating of enamel, which is an opaque glass, is very good, but it is expensive, and it chips easily. *Russia iron* is iron coated with a layer of Fe_3O_4 , a black oxide, made by blowing steam on the hot metal, and it resists rusting for years. It is used for decorative iron work. *Copper steel.* The addition of one half of one per cent of copper greatly increases the rust resistance of iron.

407. Compounds of iron. Iron forms two different ions, Fe^{++} , ferrous, and Fe^{+++} , ferric. Ferrous compounds are usually light green in color and are rather strong reducing agents, going over into the more stable ferric compounds. Ferric compounds appear red, or yellow, due to colloidal $\text{Fe}(\text{OH})_3$ formed by hydrolysis of the ferric compound.

Ferrous ion can be changed to ferric ion by oxidation; as, for instance,



where we notice $2 \text{Fe}^{++} \xrightarrow{-2e} \text{Fe}_2^{+++}$

the 2 Fe^{++} losing 2 electrons to become 2 Fe^{+++} , while the O^0 gains 2 electrons to become O^{--} . That is, the Fe^{++} is oxidized to Fe^{+++} .

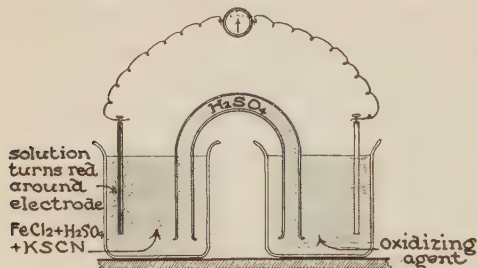
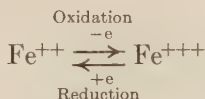


FIG. 262. Oxidation by loss of electrons.

Now since only electrons are needed to bring about this change, it should be possible to do it without actual contact between oxidizing agent and ferrous compound, and this can be done as follows:

arrange the apparatus and solutions as shown in the diagram, using potassium sulphocyanide, KSCN, to test for Fe^{+++} . In half an hour or so a red coloration around the electrode shows the change from Fe^{++} to Fe^{+++} , showing that something coming from or going into the electrode over the wire is causing the oxidation; that is, electrons are leaving the Fe^{++} , as can be shown by putting a galvanometer in the circuit. Reduction of Fe^{+++} to Fe^{++} can be accomplished in a similar way by gain of electrons.

To sum up:



408. Oxides of iron. There are three oxides, FeO , Fe_2O_3 , and Fe_3O_4 . FeO is a black powder easily oxidized. Fe_2O_3 is the common oxide, known as hematite, rouge, crocus, Venetian red. When FeSO_4 , or some of the minerals of iron, like limonite, are heated, various shades of pigment are obtained, due to different mixtures of oxides, all of which are used in paints. These are known under the following names:

COLOR	NAME
red	Venetian red Indian red Tuscan red
yellow	yellow ocher { French ocher English ocher raw sienna — Italian sienna
brown	burnt sienna raw umber burnt umber metallic brown mineral brown Vandyke brown

Fe_3O_4 , called the *magnetic oxide of iron*, is a black lustrous compound found in nature, and known as lodestone or magnetite. It is the scale that flies from the iron when it is hammered by the blacksmith.

409. Ferric hydroxide, $\text{Fe}(\text{OH})_3$, is a red compound, exceedingly insoluble in water, but forming colloidal solutions that look like wine. When dialyzed, it is used in medicine, as an antidote for arsenic poisoning, and other purposes. It is called dialyzed iron.

410. Ferric chloride, FeCl_3 , is made by passing chlorine into ferrous chloride solution and is a yellow, deliquescent substance which hydrolyzes easily:



so that a given sample of FeCl_3 will show a different analysis at different times. It is never FeCl_3 , but consists of varying amounts of FeCl_3 , $\text{Fe}(\text{OH})_3$, and HCl . In solution it is yellow to red, depending upon the amount of colloidal $\text{Fe}(\text{OH})_3$ present.

411. Ferrous bicarbonate, $\text{Fe}(\text{HCO}_3)_2$, is found in all natural waters, where it is probably formed by the solution of natural iron compounds by $\text{CO}_2 + \text{H}_2\text{O}$. It forms deposits in vessels, like boiler scale, and makes clothing yellow when washed, being oxidized to $\text{Fe}(\text{OH})_3$:



Waters containing much $\text{Fe}(\text{HCO}_3)_2$ are known as chalybeate waters.

412. Ferrous sulphate, FeSO_4 , known as green vitriol, or “copperas,” is a by-product of the steel industry. Iron pickled in sulphuric acid before plating forms FeSO_4 , and this spent acid is used with pieces of waste iron and steel to form light green crystals of ferrous sulphate. It is used for a multitude of purposes, among which may be mentioned tanning, water purification, fungicide, shoe blacking, paper, and ink. Ordinary writing ink is chiefly ferrous tannate, a colorless compound, which on exposure to air changes to ferric tannate, a black insoluble compound.

413. Pyrites, FeS_2 , is a natural compound. It is often called "fool's gold," because it resembles gold in color and luster. It is a very common mineral, found with coal and in nearly all rocks. It is this substance which causes the red ash of some coals (Fe_2O_3 formed on burning), and it is the source of most of the troubles of iron smelters, for to it is due the presence of sulphur in coke and therefore in cast iron and steel.

414. Potassium ferricyanide, $\text{K}_3\text{Fe}(\text{CN})_6$, known also as "red prussiate of potash," is used chiefly in making blueprint paper. With ferric salts it forms a brown compound, soluble in water, but with ferrous salts we get a deep blue precipitate known as Turnbull's blue. This is what gives the color to blue prints. Blueprint paper has on its surface essentially a ferric salt and a reducing agent mixed with $\text{K}_3\text{Fe}(\text{CN})_6$. The reaction is brought about by actinic light, present in the sun's rays and in other

white or violet light. To make a blue print, the paper is exposed to the light, under the negative or detail drawing on tracing linen. Wherever, on the negative, there are dark opaque lines, no light will penetrate, so that no chemical change will take place. Only where the light strikes will there be a reduction of ferric to ferrous salt. The print is then washed in water. Now the ferrous salt combines with $\text{K}_3\text{Fe}(\text{CN})_6$ to form Turnbull's blue, while the ferric salt, left unchanged, forms the soluble brown compound which is washed off.

415. Prussian or Berlin blue. This is formed by the interaction of a soluble ferric salt with potassium ferrocyanide, $\text{K}_4\text{Fe}(\text{CN})_6$,

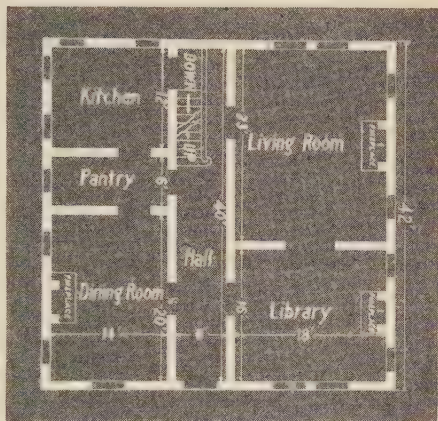


FIG. 263. Blue print.

also known as "yellow prussiate of potash." It is used as a pigment in paints, inks, and laundry bluing. With bases it forms $\text{Fe}(\text{OH})_3$, which means that if clothes are not freed from soap before bluing with Prussian blue, they will get rusty.

SUPPLEMENTARY READING

Waldemar Kaempffert, *A Popular History of American Invention* (Charles Scribner's Sons, 1924), vol. II, p. 3.

Floyd L. Darrow, *The Story of Chemistry* (Bobbs-Merrill Co., 1927), p. 329.

QUESTIONS

1. State some of the properties of iron which make it the most important metal, and give one use dependent upon each property.

2. Why is sand usually brown?

3. Name two ores of iron, and tell where each is mined.

4. Why is iron so cheap?

5. What is the chief gangue of iron ore? How is it removed?

6. State two uses of coke in the blast-furnace charge.

7. What is pig iron? Look up in the daily paper (financial sheet) the quotation for pig iron and for steel.

8. Make a diagram of the blast furnace, labeling important parts, and explain the function of each part.

9. Write an equation for the reduction of iron ore.

10. What is "spiegeliesen"? ferromanganese?

11. Under what condition would limestone not be added to the blast-furnace charge?

12. What is "malleable iron"? In what way is it superior to cast iron?

13. What is the best property of cast iron? Name an object made of cast iron because of this property.

14. What property of cast iron is a distinct disadvantage? Name an article which you have found to be made of cast iron and which, in your opinion, should have been made of some other iron alloy.

15. Name as many articles as you can (not in the text) which are made of cast iron.

16. What relation has the type of furnace used to the purity of wrought iron?

17. What is the difference in properties between wrought iron and cast iron?

18. Why can steel be made either from cast iron or from wrought iron? Which makes the better steel?
19. Why is air blown into the converter?
20. Why is open-hearth steel better than Bessemer?
21. Why is electrically treated steel superior to the ordinary variety?
22. What is the chief property of steel? Name one article made of steel because of this property.
23. Steel has a wide variety of properties. Arrange these in a vertical column, and next to each place the name of one article which makes use of that property.
24. Explain, in your own words, how to shape and temper a razor blade.
25. Show how the presence of manganese makes it possible to produce structural steel.
26. Select three alloy steels, each of which possesses a property not found in ordinary carbon steel, and name an article made of the alloy for that reason.
27. Why is water and oxygen necessary for rust formation?
28. What is enameled ware?
29. Tell how to remove rust stains.
30. How may rust stains be produced in washing clothes?
31. What is yellow ocher; burnt sienna; raw umber; fool's gold?
32. Explain the chemistry of blue printing.
33. Of what alloy of iron should each of the following be made: radiator; stove; piston rod; wire rope; "I" beam; post; sheet iron; gas pipe; spring; knife; wired glass?
34. In the Bessemer process, explain why *molten* cast iron is used; why is carbon first removed and then added?
35. Why is a firebrick lining necessary in the blast furnace?
36. By what physical property may one distinguish a brass plated door knob from solid brass?
37. Write an equation to show how a ferric compound may be changed to ferrous; ferrous to ferric.
38. Commercial ferrous sulphate contains some ferric salt. Explain its presence.
39. State the essential differences in composition between cast iron, wrought iron, and steel.
40. There is plenty of iron in ordinary earth. Why is this not extracted?
41. What chemical principles are involved in the change from cast iron to wrought iron; cast iron to steel?
42. What substances beside coke might be used in the reduction of iron ore? Why is coke preferred?

43. Why is it necessary to manufacture cast iron, first, rather than to make steel directly from the ore?
44. What prevents the molten cast iron from running out at the bottom of the Bessemer converter?
45. How does the presence of sulphur and phosphorus in steel affect its properties?
46. Name as many substances as you can which will change Fe^{++} to Fe^{+++} and vice versa.
47. Show by equation how slag is formed in the blast furnace. In what way does this resemble glass? How does it differ?
48. Write equations to show the action of iron on hydrochloric acid; dilute sulphuric; concentrated hot sulphuric; nitric acid. (Remember that nitric and hot concentrated sulphuric acids are oxidizing agents.)
49. How many tons of coke is it necessary to add to 100 kilograms of iron ore, 60 per cent pure, in the blast furnace? What volume of carbon monoxide is produced?
50. Explain the red color of ferric compounds.

OPTIONAL QUESTIONS

51. In two columns, one headed iron, the other, any other metal you please, make a list of as many properties of their alloys as you can. What conclusion do you draw from such a table?
52. How might you distinguish, by chemical tests, Fe^{++} from Fe^{+++} ?
53. From your knowledge of the electron structure of matter, what must be the essential difference between Fe^{++} and Fe^{+++} ? How would you account for the different properties of ferrous and ferric compounds?
54. What is the essential difference between cuprous (Cu^+) and cupric (Cu^{++}) ions? (Consider ferric and ferrous ions.)

CHAPTER XXIX

ELECTRICITY AS APPLIED TO CHEMISTRY

416. It is a far cry from the electromagnet of Faraday to the huge dynamos of our own day; yet we must not forget that Faraday's original discovery has made possible not only the use of electricity for transportation and lighting purposes, but its use in a number of highly important chemical industries. Where water power is available and electrical power, therefore, relatively cheap, as at Niagara Falls, there electrochemical industries flourish; and those of you who by this time have seen the electrolysis of water as a demonstration experiment will have little difficulty in finding sound reasons why electrochemical methods should flourish, provided electrical power can be obtained at a reasonable price.

Electricity in chemistry is used in two ways: either to produce chemical changes, to decompose substances (as in the electrolysis of water, in the preparation of sodium hydroxide and of aluminum), or as a source of very intense heat — much above what can be obtained from the burning of any of the common fuels (as, for example, in the preparation of calcium carbide and carborundum). We shall first discuss those industries which depend upon electrolysis.

417. The electrolysis of water, which has already been discussed, is often used as a commercial source of hydrogen, which, in turn, is so largely used nowadays in hardening vegetable oils. The electrodes are made of sheet iron; and as many as 200 cells may be found in one room of the factory.

418. Electroplating. Nickel, silver, and gold plating are brought about by electrolysis. The object to be plated is made

the cathode, so that the metallic ion, Ni^{++} , or Ag^+ , or Au^{+++} , gains electrons, clinging to the cathode as a thin coating or plate. The action would become slower and slower due to discharge of more and more ions, if it were not for the fact that the anode is made of pure nickel, silver, or gold. In this case for each atom of metal deposited on the cathode, one atom is dissolved off the anode to replace it, thus keeping constant the number of ions in

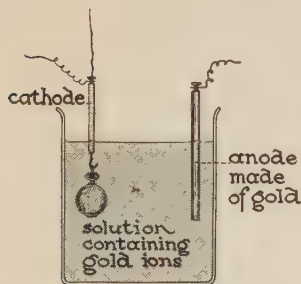


FIG. 264. Gold plating.

solution. Let us apply this to gold plating, using gold chloride (AuCl_3) as electrolyte. The anode is made of gold, the cathode is, let us say, a watch case, to be gold plated. The AuCl_3 ionizes into Au^{+++} and 3Cl^- , on solution. The gold atoms of the anode lose electrons, becoming Au^{+++} (gold ions). The gold ions migrate to the cathode, where they receive three

electrons, become metallic gold, and deposit on the cathode, which is the watch case. As a result of all this one atom of gold has been dissolved off the anode and deposited on the cathode, while the solution still remains of the same strength. This makes the process uniform, so that one can predict how much gold will be deposited on the cathode per hour. There are many practical points in plating which we cannot discuss here, except that rapid deposition of metal (*i.e.*, from solutions of high concentration, or by great electrical pressure or voltage) gives colloidal metal, black, which will not stick. Shiny metallic deposits are crystalline and take longer to form. To prevent rapid deposition, therefore, it is customary to use substances such as certain organic acids or gelatin or gum arabic, as protective colloids.

Most of the copper used, the so-called electrolytic copper, is refined by this method. The impure copper obtained from smelting is cast into anode plates, which are suspended in an electrolyte of CuSO_4 solution. Pure copper is dissolved off the anode and

deposited on the cathode, while the impurities in the anode drop to the bottom.

419. Electrotyping. This is done by electrolysis. The type is set up and an impression made in wax. This is dusted over with graphite, a fair electrical conductor, and hung in a solution of copper sulphate on the cathode. Copper is deposited on the graphite, and when a thin sheet is formed, it is stripped and backed with molten lead. This makes the plate with which the actual printing is done. It lasts longer and is cheaper than type. Photographs are printed in books and newspapers in a similar way; the plate



FIG. 265. Phonograph record made from electrottype.

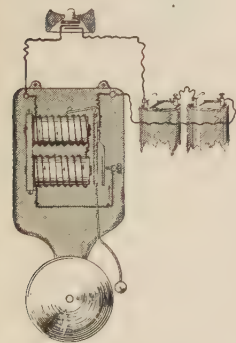


FIG. 266. Electric bell.

being made from the photographic negative. Phonograph records are made on a wax original, but the large number of hard rubber duplicates that are sold are reproduced from an electrottype. This electrottype is made by hanging the original record covered with graphite on the cathode in a solution of zinc sulphate. In this way a zinc plate is made and can be used to make impressions in soft rubber, which is then vulcanized.

420. Electrical cells and batteries. The "primary cell" consists of a strip of zinc and a strip of copper immersed in a solution of sulphuric acid. When the circuit is closed, Zn^0 loses two electrons to the copper, becoming Zn^{++} ; while the 2H^+ receive 2 electrons from the Cu, and become H_2^0 . That is, metallic zinc is dissolved and H_2 liberated;

and this action will continue until all the zinc or all the H^+ is used up. The liberated electrons perform work, as, for example, in our

diagram ringing a bell. Our *dry battery* is a modification of this cell, consisting of zinc and carbon electrodes and a paste of moist ammonium chloride (sal ammoniac), carbon, and manganese dioxide, sealed up to prevent drying. The zinc loses electrons while NH_4^+ gains electrons at the carbon electrode. The reaction is:

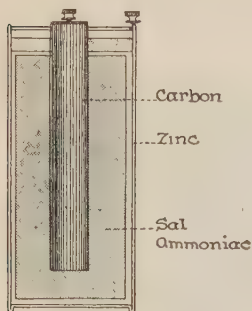
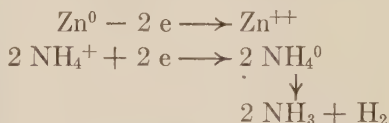


FIG. 267. A dry cell.



The MnO_2 combines with H_2 to form water. It will now be understood that a primary cell of this kind requires only a cathode of some substance which can give up electrons (reacts chemically, such as metal + acid) and an anode, less active, which will receive electrons; in other words, an electrolyte and two materials, one of which tends to go into solution while the other does not. Of course both materials must be electrical conductors.

421. The storage battery is an example of several secondary cells. The ordinary 6-volt battery contains three cells, one of which is shown diagrammatically. (The actual cell has more plates to increase its power.) Each cell has a lead plate and a lead grid holding lead peroxide, both immersed in sulphuric acid.

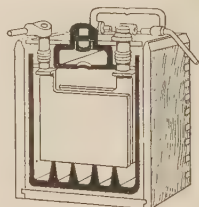
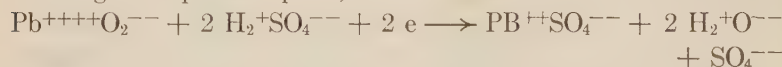


FIG. 268. The storage battery.

In discharging, the cell acts as follows:

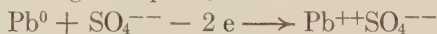
At the grid or positive plate,



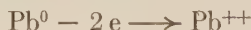
or, in short,



While, at the minus or negative plate,



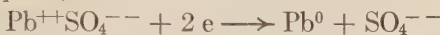
that is,



At one plate, then, electrons are lost, at the other gained; thus producing the electric current. (Commercial storage batteries have + and - marked correctly for charging, but we must remember these poles are incorrectly marked for discharging.)

During the charging process, these changes are reversed:

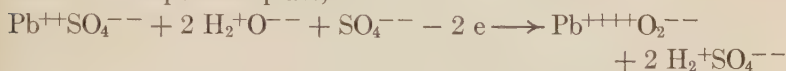
At the negative plate,



that is,



While at the positive plate,



or

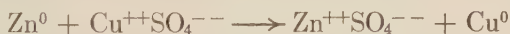


These changes restore original conditions; that is, we have again lead and lead peroxide on the plates and sulphuric acid in solution.

422. The electromotive series. It has been noticed that some elements tend to lose electrons easily while others gain them. Zinc, for instance, easily goes into solution, that is, loses electrons; also sodium, — while hydrogen and copper easily gain electrons. In other words, zinc replaces hydrogen:



The zinc has lost two electrons, while the 2H^+ have gained two electrons to form H_2^0 . Similarly



The Zn^0 has lost two electrons to the Cu^{++} . In this way copper replaces silver:



and zinc likewise replaces silver:



whereas silver does not replace Zn or Cu, nor does Cu replace Zn. If we place these 3 metals in the order of Zn, Cu, Ag, the arrangement is such that any one replaces those after it; that is, tends to

give up electrons to it. If we arrange the other common metals in this order, we get a list called the **ELECTROMOTIVE SERIES** of the metals:

K	Each element will give up electrons to any one that
Na	follows it; that is, K replaces any metal below it; Zn
Ca	replaces those below it; while Au replaces none of
Mg	the metals listed. This is a very useful table. For
Al	instance, chemical activity decreases going down. All
Zn	those above H corrode easily, those below do not;
Fe	which explains why metallic Na, Ca, etc., are not found
Ni	free in nature, while those below H are found free.
Sn	The relative positions of Zn and Sn (tin) explain why
Pb	Zn makes a better protective covering for iron than
Hydrogen	tin. Let us assume that we have iron and tin in
Cu	water. The tendency there would be for the iron to
Hg	lose electrons and go into solution; that is, iron rusts.
Ag	On the other hand, suppose we have iron and zinc in
Pt	water; in this case the zinc tends to lose electrons,
Au	while the iron does not (therefore no rusting).

423. Sodium hydroxide (caustic soda) is now made very largely by electrolyzing a concentrated brine solution:



the difficulty in the process consisting very largely in preventing the chlorine from coming in contact with the caustic soda and forming sodium hypochlorite. To prevent this, Castner devised a cell consisting of three compartments, the walls of which dip into a trough of mercury. The two end compartments contain solutions of brine and anodes of carbon; the central compartment contains water and an iron cathode. The chlorine is discharged at the anodes and the sodium combines with the mercury, forming sodium amalgam, which, by rocking the apparatus, finds its way into the central compartment, and there reacts with the water, forming sodium hydroxide.

A modification of the Castner cell was devised by Nelson, and the Nelson cell is now very largely used. Here the brine is in a steel container in which are immersed the anodes made of carbon. An asbestos cloth surrounding the anodes separates the latter from the cathode, made of steel. The asbestos prevents, for a time, the sodium hydroxide liberated at the cathode from mixing with the chlorine given off at the anode. The separation is more effectively accomplished by running in from the top a slow stream of sodium chloride and allowing the contents to drip from the bottom. This gives a mixture of sodium chloride and sodium hydroxide, and the former is removed by fractional crystallization.

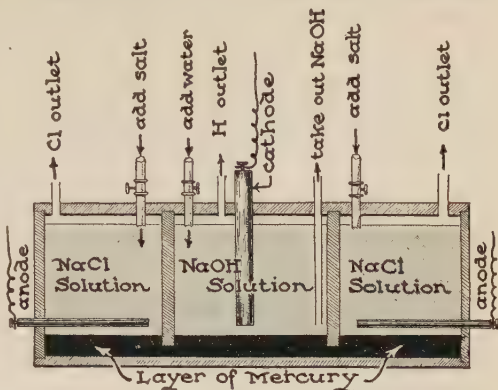


FIG. 269. The Castner process for the manufacture of sodium hydroxide.

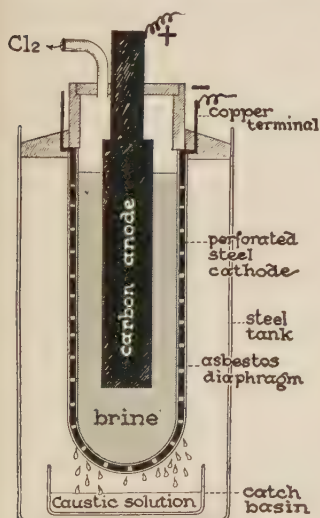


FIG. 270. The Nelson cell for the manufacture of sodium hydroxide.

Sodium hydroxide is used in the manufacture of soap, the refining of petroleum, the making of paper, and the manufacture of artificial silk.

The process just outlined also serves, of course, as an excellent industrial source of chlorine and hydrogen. Usually the demand for sodium hydroxide is greater than the demand for either chlorine or hydrogen. However, during the late war, the United States Government manufactured at

Edgewood Arsenal enormous quantities of chlorine, using the Nelson cell arrangement.

424. Aluminum. One of the greatest triumphs of the application of electricity to chemistry is found in the manufacture of aluminum. Though widely scattered in nature, in the form of various clays and other compounds, the element itself remained a chemical curiosity until Hall in 1886 devised an ingenious method for its extraction; and thereby lowered the cost of the metal from about \$100.00 a pound to about 25¢ a pound.

The first method which naturally suggests itself is to reduce the oxide of aluminum with carbon just as we may reduce iron oxide or copper oxide; but when this is tried, not aluminum but aluminum carbide, Al_4C_3 , is obtained. Wöhler, years ago, did obtain small quantities of the metal by heating the chloride with sodium (a powerful reducing agent).

To understand the principle employed by Hall, let us, for a moment, revert back to the electrolysis of sodium chloride. If the attempt be made to pass a current through solid sodium chloride, nothing will happen. We say that solid sodium chloride is a non-conductor. However, by first finding a suitable solvent for the salt, such as water, there results a solution which does allow a current to flow through it.

Now let us apply this conception to the electrolysis of an aluminum compound. We select for our purpose the mineral bauxite, which, when properly purified, is essentially aluminum oxide, Al_2O_3 . A current can no more decompose aluminum oxide than it can sodium chloride; but if we first dissolve the aluminum oxide in molten cryolite (a complex sodium aluminum fluoride found in nature), the solution becomes a conductor and the Al_2O_3 is decomposed thus:



The discovery of a suitable solvent is the key to Hall's work on aluminum.

In actual practice, cryolite is put into a rectangular iron vessel lined on the inside with carbon, which is the cathode. Carbon plates, acting as the anode, are lowered, the current turned on, and the heat generated melts the cryolite. At this point the aluminum oxide is added; this dissolves in the cryolite and is next decomposed by the current, as already explained. The aluminum, being positive, collects at the cathode, where it is removed; and the oxygen, being negative, collects at the anode.

Its lightness, its comparative resistance to rust, and the fact that it is an excellent conductor of heat, has made aluminum an article very much sought after. It finds use in kitchen utensils; as aluminum foil instead of tin foil; as an electrical conductor, particularly over great distances with few available supports; in the building of airplanes

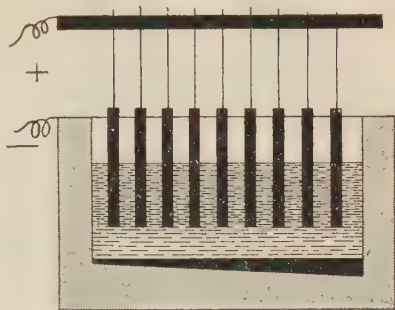
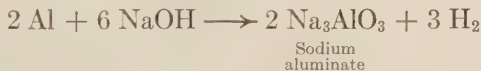


FIG 271. The extraction of aluminum.

(in the form of its alloys, such as magnalium, which contains magnesium); as aluminum bronze (which contains copper and which is stronger than bronze); as nickeloxy (which contains copper and nickel) for castings; in removing inclosed bubbles of gas ("blowholes") from molten steel; etc. The fact that alkalis attack and dissolve aluminum:



makes it important that aluminum utensils be not washed with strong soaps and alkalies.

But perhaps the most outstanding use for aluminum is in the *thermite process*, which makes use of its powerful reducing property and also of the enormous amount of heat liberated in the reaction.

Many of the most stable oxides can be decomposed by heating them with aluminum. For example, iron oxide and aluminum



FIG. 272. Welding a rail by the thermite process.

need merely be ignited, when a violent reaction takes place :



and an enormous amount of heat is generated. In fact, the heat is so intense that the iron formed is in a fluid condition. This fact is extensively made use of in rapidly repairing broken machinery. The casting, surrounded by a mold of sand, is heated to redness by a

torch at the points to be joined. The molten iron formed in the funnel-shaped crucible by the interaction of the aluminum and iron oxide flows into the mold. In this way a weld is accomplished.

425. So far we have discussed electricity as a chemical agent in the sense of decomposing substances. We shall now discuss the use of electricity in industry as a means of acquiring very high temperatures.

426. Calcium carbide. One of the earliest of such uses was in the manufacture of calcium carbide, CaC_2 , using the *arc* type of furnace. The method in brief is to mix quicklime and coke and heat the mixture to a very high temperature :



The furnace in its simplest form consists of two blocks of lime with central cavity for the crucible containing the material to be used and horizontal cavities for the graphite electrodes. The electrodes are first made to touch one another, the current is turned on, and the electrodes

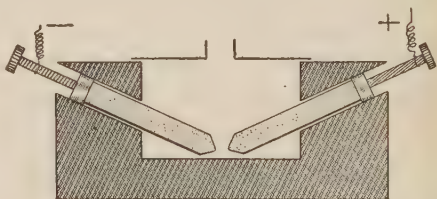


FIG. 273. Making calcium carbide in an arc furnace.

are gradually parted, when an *arc* is formed between the two; and this is the source of the intense heat.

The carbide so produced finds an extensive market in the preparation of acetylene and of calcium cyanamide.

427. Carborundum (silicon carbide), SiC , is another product of the electric furnace resulting from a high temperature reaction between silica and coke:



Here the furnace used is usually of the resistance rather than the arc type; which means that instead of the electrodes being quite near one another, forming an arc between them, they are fairly far apart and have a core of coke in between them; the coke setting up a very high resistance to

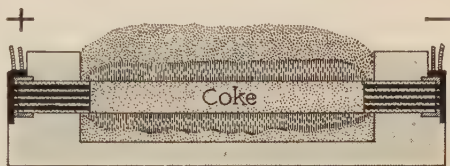


FIG. 274. Making carborundum in a resistance furnace.

the passage of the current, and thereby generating much heat, which brings about the reaction. The carborundum so produced (a discovery of Acheson) is almost as hard as diamond and finds, therefore, very extensive use as an abrasive.

428. Graphite is now very largely made in a furnace similar to that used in the manufacture of carborundum. Here the raw material is either coke or anthracite coal; and the high temperature of the furnace converts much of this coke or coal to graphite. You will recognize here a means of converting an amorphous variety of carbon into one of its crystalline forms (chap. xii). Lead pencils are made from graphite and clay mixed in varying proportions. The softer pencils have more graphite. Graphite is unique among the forms of carbon in being a conductor of electricity; and so finds extensive use in the manufacture of electrodes. Finely divided graphite may be suspended in water (aquadag) or in oil (oildag) and used as a lubricant, particularly at high temperatures.

429. Refining of steel. Bessemer steel is often further refined or purified by an electrical method (chap. xxviii). The main advantage of the electrical process over the open-hearth is that the control over the product formed, good as it is in the open-hearth process, is even better in the electrical process; and there is a total elimination of "blowholes." Sulphur and phosphorus, two of the most objectionable impurities, are almost completely eliminated.

SUPPLEMENTARY READING

E. E. Slosson, *Creative Chemistry* (Century Co., 1919), chap. xiii.

S. A. Arrhenius, *Chemistry in Modern Life* (D. Van Nostrand Co., 1925), chap. x.

H. E. Howe, *Chemistry in Industry* (Chemical Foundation, 1924), vol. I, chaps. vi and vii; vol. II, chaps. vii and viii.

QUESTIONS

1. Write as many equations as you know to represent the preparation of (a) sodium hydroxide, (b) chlorine, (c) hydrogen.

2. What is the outstanding difficulty in the preparation of sodium hydroxide by the electrolytic method?

3. Name the various uses for the products obtained when brine is electrolyzed.

4. How many tons of sodium chloride, 99 per cent pure, would be necessary to obtain 100 tons of sodium hydroxide?

5. Write an equation to represent one important use for sodium hydroxide and state what the use is.

6. What quantitative relationship is brought out when water is electrolyzed? How would you identify the products produced? What would happen to these products if brought together and sparked?

7. Why is a small quantity of sulphuric acid or sodium hydroxide added to water before the latter is electrolyzed?

8. How much (a) oxygen, (b) hydrogen are obtained when 100 grams of water are electrolyzed? What volume do these gases occupy?

9. Write equations to represent the reduction of two common oxides. What method might this suggest for the preparation of aluminum? Why cannot such a method be used?

10. Explain the function of (a) the bauxite, (b) the cryolite, (c) the electric current in the preparation of aluminum.

11. Taking into consideration the place of aluminum in the electro-motive series, write equations between aluminum and (1) hydrochloric acid, (2) sulphuric acid, dilute and hot concentrated, (3) potassium hydroxide, (4) lead nitrate, (5) magnesium sulphate.

12. Give the equation representing the reaction in the thermite process, but use an oxide other than iron. What property does the aluminum exhibit in this reaction? Name as many substances as you know other than aluminum which show this property to a greater or less degree, and in each case illustrate with an equation.

13. Give as many uses for aluminum as you can and in each case name the property of the metal associated with its use.

14. How many tons of aluminum can be obtained from 100 tons of bauxite, 95 per cent pure?

15. State the principle of the method employed in electroplating.

16. How would you electroplate (a) an iron gas stove with nickel, (b) a spoon with silver, (c) a watch case with gold.

17. Describe how the type of a book is often prepared.

18. Describe briefly the two types of electric furnaces in common use. What are the essential differences?

19. Starting with limestone, show how calcium carbide may be prepared. Give equations.

20. Of what value are two of the products obtained from calcium carbide?

21. What product might be formed if silicon and carbon were very strongly heated? Why is this method not used to prepare the product? (Consider cost.)

22. How can amorphous carbon be changed into each of its two crystalline forms?

23. Give the chemical names for the following: (a) caustic soda, (b) water, (c) pure bauxite, (d) brine, (e) coke, (f) carborundum, (g) silica.

24. What advantages have electrical methods over other methods for preparing compounds?

25. In the Hall process why are carbon anodes used rather than iron?

26. What property of aluminum accounts for its use in the purification of steel?

27. Outline, with diagram, the commercial process of electrolytic copper refining.

28. Draw two diagrams showing an ordinary storage battery discharging and being charged. Label parts, electrons in red, using arrows to show the migration of electrons in each case.

29. Draw a diagram of a dry battery and explain its action.

30. Suppose five iron cans are plated with zinc, tin, aluminum,

nickel, and copper, respectively, and exposed to moist air. In what order will corrosion, if any, take place?

31. Why is not thermite used as fuel in the home?

32. In the reaction $\text{Cu} + 2 \text{AgNO}_3 \longrightarrow \text{Cu}(\text{NO}_3)_2 + 2 \text{Ag}$ show, by rewriting this electronically, that we are dealing with an oxidation-reduction reaction. Which is the oxidizing and which is the reducing agent?

OPTIONAL QUESTIONS

33. How would you deposit a plating of brass on an iron door knob?

34. Explain with the aid of a diagram how metallic sodium or potassium is extracted from the fused hydroxide.

CHAPTER XXX

CARBON COMPOUNDS

430. "The Chemistry of the Carbon Compounds" is commonly known as *Organic Chemistry*, because for a long time it was supposed that such compounds were in some way intimately connected with the life process. It was supposed that such compounds could never be prepared artificially by man. Then in 1828 this theory was completely upset by the chemist Wöhler, who showed how to prepare urea in his laboratory. Now if there is any one compound that can be called "organic," that distinction certainly belongs to urea, because it is the principal waste product in the body due to the decomposition of proteins.

Wöhler's work was soon followed by the synthesis, in the laboratory, of many other compounds hitherto regarded as typically organic. Gradually but surely the significance attached to the name organic disappeared; and though we still retain the name "organic chemistry," we, in reality, mean nothing more than the chemistry of the carbon compounds.

431. Carbon compounds are many and important. There are some 225,000 compounds of carbon which have so far been isolated, analyzed, and classified. If the compounds of all the elements other than carbon be added together, the total does not exceed 26,000. How is it that the compounds of one element should outnumber the compounds of the other 91 elements in the ratio



Brown Bros.

FIG. 275. Wöhler.

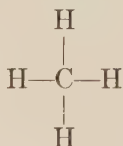
of about 10 to 1? The answer to this question is to be found in the unique property which the atoms of carbon possess of combining with one another to form chainlike structures. This will become clear as we proceed with the chapter.

432. Importance of organic chemistry. The chief constituents of the cell — the fats, the proteins, and the carbohydrates — are all carbon compounds; so that organic chemistry plays a very important rôle in the study of life phenomena and foods. And in the industries we find many of our best-known substances to be compounds of carbon. Let us name a few of them. There are starch, sugar, paper, rayon, soap, explosives, various anaesthetics, disinfectants, antiseptics, dyes, drugs, waxes, perfumes, camphor, rubber, gasoline, glycerol, etc.

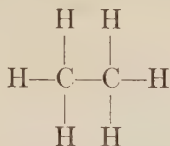
THE HYDROCARBONS

433. Among the simplest groups of compounds in organic chemistry are the hydrocarbons, which, as the name implies, consist of the elements hydrogen and carbon. The simplest compound of the hydrocarbon series is *methane*, or marsh gas, with the formula CH_4 . The decomposition of vegetable and animal matter gives rise to this gas. One of the gases arising from marshes is methane; hence the name "marsh gas." It is also one of the gases produced in intestinal putrefaction. It forms a large percentage of the constituents found in natural gas (80 per cent and above) and coal gas (30–40 per cent). Fires and explosions in coal mines are mainly due to mixtures of methane and air (hence the name "fire-damp" for methane).

If we remember that the valence of carbon is 4 and that of hydrogen 1, we can represent the formula of methane as



where each bond represents a valency of 1. A closely related compound to methane is *ethane*, C_2H_6 , which may be represented as



You will notice that the difference between CH_4 and C_2H_6 is CH_2 . By starting with methane, the chemist has been able to build up a whole series of compounds of carbon and hydrogen, where the difference between any two members is just CH_2 . Thus

methane, CH_4
ethane, C_2H_6
propane, C_3H_8
butane, C_4H_{10}
pentane, C_5H_{12} , etc.

Such a series is called a "homologous series"; and you will see that it may be given the general algebraic expression C_nH_{2n+2} . (If, for example, $n = 1$, then C_nH_{2n+2} becomes C_1H_{2+2} , which is CH_4 .) Owing to the fact that this series of compounds is relatively stable, they are known as the *paraffins*, which means "little affinity" (for other compounds).

434. Petroleum. These paraffins are of enormous practical importance, since various mixtures of them make up, to a very large extent, the product which we know as "petroleum." The Indians in Western Pennsylvania first discovered oil floating on surface waters. By them it was used as a remedy for all physical ills. In the middle of the last century it occurred to a Colonel Drake that since oil came to the surface of springs, it was probably present in much larger quantities beneath the earth's surface. He thereupon proceeded to drill a well near Oil Creek, Pennsylvania, and, before he had dug 100 feet, oil came to the surface in such quantities that it could not be collected.

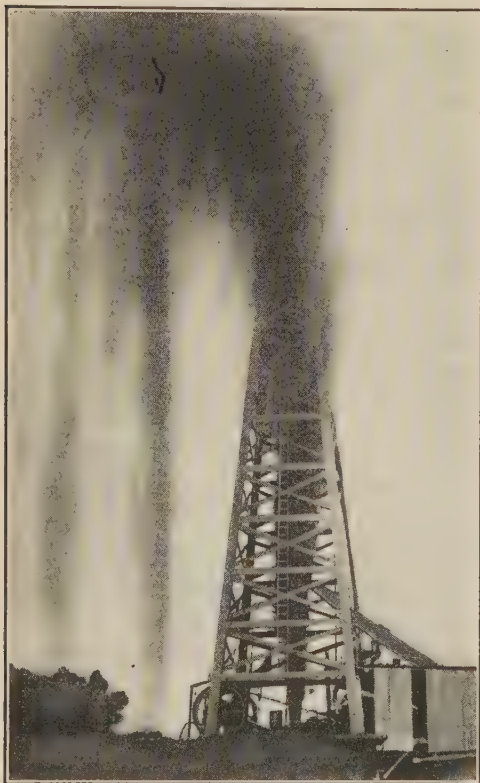
Petroleum is found not only in America (Oklahoma, Pennsylvania, California, and Texas), but in Mexico, Roumania, Russia, and Persia. The various important petroleum products (such as naphtha, gasoline, kerosene, gas oil, lubricating oil, cylinder oil,

vaseline, etc.) are obtained from petroleum by a process of *fractional distillation*.

The principle of this process is based upon the fact that when a mixture of two or more liquids, having different boiling points, is distilled, the liquid with the lower boiling-point distils over first. When petroleum is distilled, the first fraction (which consists of everything that distils over up to $150^{\circ}\text{C}.$) includes petroleum ether, benzine, gasoline, and ligroin; the fraction from 150° to 300° includes the illuminating oils, of which kerosene is the best example;

the fraction from 300° up includes spindle, machine, mineral, and cylinder oils.

In addition, many products of commercial value, such as vaseline and paraffin, are obtained; and the tar residue which remains in the still is used in road-making, artificial asphalt, roofing, etc.



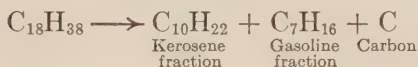
Ewing Galloway

FIG. 276. An oil well.

If the temperature is high enough, we obtain petroleum coke in the place of tar. Due to its high purity, this coke finds extensive use in the manufacture of electrodes.

Commercially, the most important product obtained from petroleum is gasoline, which is a mixture of hydrocarbons of relatively low molecular weight, such as pentane C_5H_{12} , hexane C_6H_{14} , heptane C_7H_{16} , etc. The process of purification consists in treating the gasoline with sulphuric acid — incidentally one of the most important uses for this acid — whereby many objectionable impurities are removed; the sulphuric acid in turn being removed by washing first with water and then with sodium hydroxide solution.

By a careful study of the effects of various temperatures and pressures upon the course of the distillation, American chemists have developed a *cracking process*, by which the more complex hydrocarbons are converted into the simpler hydrocarbons; for example,



This represents an important advance in methods by which the amount of gasoline obtained from petroleum may be increased.

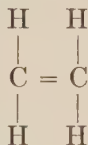
Times have changed. Less than thirty years ago kerosene cost more than gasoline; the latter, in fact, was regarded as little more than a nuisance. To-day it would be hard to conceive of substances more valuable in commerce than gasoline. Wherever minute quantities of this fuel can be found it is carefully extracted. Even the small amount which occurs in natural gas is extracted and recovered, and sold as "casinghead" gasoline.

A very promising method of manufacturing petroleum and making gasoline ("synthetic gasoline," as it has come to be called) has been developed in Germany. Low-grade coal, usually lignite, is finely powdered, heated to about $500^\circ C.$, and hydrogen is introduced, the pressure being kept at about 200 atmospheres. The hydrogen and the carbon (from the coal) combine, under these conditions, forming a mixture of hydrocarbons resembling petro-

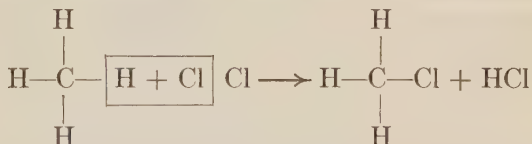
leum. Gasoline is recovered from this mixture by the usual fractional distillation method.

Commenting on such achievements as the one just mentioned, a writer notes: "The characteristic feature of chemical production is that it is gradually making the human economic system independent of the natural distribution of the supply of raw materials"; which, translated into simpler language, means that the chemist in his laboratory very often successfully imitates nature, and sometimes even surpasses her.

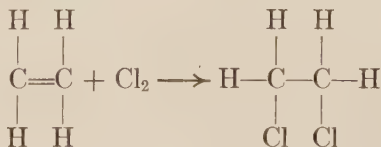
435. Ethylene, C_2H_4 , is another hydrocarbon which, like methane, is the starting point for a whole series of hydrocarbons known as the *olefines*, with the general formula C_nH_{2n} . Keeping in mind that the valence of carbon is 4, the structure of ethylene may be represented as



When chlorine, Cl_2 , is brought in contact with methane under suitable conditions, a reaction will take place; for example,

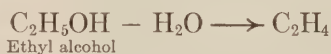


Notice that what happens is that a chlorine atom *takes the place of* a hydrogen atom. When chlorine is brought in contact with ethylene, however, a certain amount of chlorine will attach itself to the ethylene molecule, without any of the hydrogen of the latter, leaving the molecule; thus



Compounds like ethylene, which can *add* atoms of chlorine (or other elements) to themselves, are known as *unsaturated* compounds; compounds like methane which cannot add atoms to their molecules, but in which atoms change place, some entering and some leaving the molecule, are known as *saturated* compounds. (A saturated compound, then, holds as many atoms as it can.)

Ethylene may be prepared by treating ethyl alcohol with a strong dehydrating agent, like phosphorus pentoxide or concentrated sulphuric acid. The dehydrating agent takes out the elements which constitute water; thus:



Ethylene is present in coal or illuminating gas to the extent of 4 to 6 per cent and is partially responsible for the luminosity of the gas. It was used during the war in the preparation of mustard gas. More recently Dr. Luckhardt has shown that it has powerful anæsthetic properties, and that it has some advantages over nitrous oxide or laughing gas, which has hitherto been so largely employed in operations.

The story of this important discovery is worth recording. In 1908 Wisconsin carnation growers began to complain that their flowers which were shipped to Chicago and kept in greenhouses, spoiled sooner than was expected. The cause was traced to the illuminating gas, which leaked out from the gas mains. It was at first supposed that the carbon monoxide in the gas acted as a poison. But the surprising discovery was made that the principal disturber was not carbon monoxide at all, but ethylene, which, as has already been pointed out, is also present. One part of ethylene in two million parts of air sufficed to close the open flowers in the course of a twelve-hour exposure.

This led Luckhardt to investigate the effect of ethylene on animals. Frogs, rabbits, and kittens were given quite large doses of it; all the effect it had was to put them into a peaceful slumber. Far from being toxic to animals, ethylene was found

to have marked anæsthetic properties. Sublimely confident in the outcome, Luckhardt inhaled some of the gas himself. The result was as gratifying as Davy's experiment a century earlier with nitrous oxide or "laughing gas." Surgeons all over the country seized upon ethylene, the new anæsthetic. It has already been used in more than 50,000 operations.



Underwood and Underwood

FIG. 277. An oxy-acetylene blowpipe in action.

Another possible use for ethylene, also uncovered quite recently, is in accelerating the ripening of fruit, and for coloring fruits such as oranges and lemons.

436. Acetylene,

C_2H_2 , is still another

hydrocarbon which forms the starting point to many other hydrocarbons, such as C_3H_4 , C_4H_6 , etc. (Notice that the difference between any two compounds is CH_2 . This is also an homologous series.) Its structure may be represented as



and, like ethylene, it is an unsaturated compound. It is best prepared by the action of calcium carbide on water :



When mixed with air in a special type of burner, acetylene burns with a very brilliant white light and is used for illuminating pur-

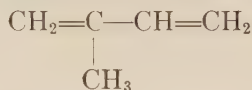
poses. When burned, it gives out a large amount of heat. This is made use of in the oxy-acetylene torch (for cutting steel, etc.), in which acetylene, supplied under pressure, is burned in the presence of oxygen. The gas is apt to explode if stored under pressure, but can be safely handled if it is first dissolved in acetone (as in "presolite" tanks).

An interesting reaction in connection with acetylene is that when passed over finely divided nickel, three molecules of it *polymerize* to form benzene :



(Polymers are substances having the same percentage composition but different molecular weight. C_2H_2 and C_6H_6 are therefore polymers.)

437. Isoprene is an unsaturated hydrocarbon of great interest. Its formula is C_5H_8 , which, if written structurally, is

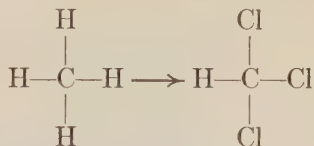


By this time you will perhaps have noticed that all unsaturated compounds have one or more double or triple bonds in the molecule. Ethylene has one double bond; acetylene has one triple bond; isoprene has two double bonds. Isoprene is one of the products obtained when caoutchouc (natural rubber) is decomposed. On the other hand, isoprene itself (in the presence of sodium, for example, which acts as a catalyst) has been converted back into a substance resembling caoutchouc, or rubber. While the synthesis of rubber on an industrial scale still awaits solution, there seems to be little doubt that the starting point for such a synthesis will prove to be isoprene.

HALOGEN DERIVATIVES OF THE HYDROCARBONS

438. Chloroform, CHCl_3 , is, as its formula indicates, a compound which contains one of the halogens, chlorine, in addition to hydro-

gen and carbon. It may be regarded as being derived from methane, in which three of the hydrogen atoms of methane are replaced by three chlorine atoms; thus:



Chloroform can actually be prepared from methane by the action of a calculated amount of chlorine, but in practice it is prepared by the action of chlorine (in the form of bleaching powder) on ethyl alcohol or acetone. It is now also made on a large scale by the reduction of carbon tetrachloride:

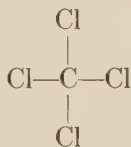


The anæsthetic properties of chloroform were discovered by Dr. Simpson in 1848. This liquid is now rarely used as an anæsthetic, for ether has largely replaced it. The advantage of ether over chloroform is that it is less dangerous and the after-effects are not so pronounced.

439. Acetone and chloroform combine to form a substance known as chloretone, which is used extensively as a hypnotic (drug which produces sleep). With concentrated nitric acid chloroform forms chloropicrin, a substance that was used in the late war as a lachrymator (tear gas). Chloroform is an excellent solvent for many organic compounds such as fats, rubber, etc.

440. *Bromoform*, CHBr_3 and *Iodoform*, CHI_3 , are compounds with properties similar to chloroform, though less pronounced.

441. **Carbon tetrachloride**, CCl_4 , may be regarded as a compound derived from methane, in which all of its four hydrogen atoms are replaced by chlorine atoms; thus:



Commercially it is made by passing chlorine into carbon disulphide, in the presence of iron as a catalyst :



It is a good solvent for gums and resins (used in vulcanizing rubber), as a constituent of many cleaning solutions (Carbena), and is used in fire extinguishers (Pyrene).

THE ALCOHOLS

442. Methyl alcohol, methanol, wood alcohol, CH_3OH , may be regarded as derived from methane, in which one of the hydrogen atoms is replaced by the hydroxyl group :

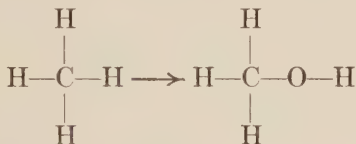


FIG. 278. Carbon tetrachloride used in a fire extinguisher.

You will notice that here we are dealing with a compound containing the elements carbon, hydrogen, and oxygen.

From *methane*, CH_4 , we get *methyl alcohol*, CH_3OH

From *ethane*, C_2H_6 , we get *ethyl alcohol*, $\text{C}_2\text{H}_5\text{OH}$

From *propane*, C_3H_8 , we get *propyl alcohol*, $\text{C}_3\text{H}_7\text{OH}$. And so on.

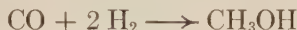
The word "alcohol" to the organic chemist means a carbon compound containing an OH group; and is not confined to methyl alcohol or ethyl (grain) alcohol.

You may notice that just as there are many hydrocarbons related to methane — C_2H_6 , C_3H_8 , etc. — so there are many alcohols related to methyl alcohol — $\text{C}_2\text{H}_5\text{OH}$, $\text{C}_3\text{H}_7\text{OH}$, etc. The consecutive members among the alcohols differ, in fact, by CH_2 , just as do the various hydrocarbons. The alcohols, therefore, also form a homologous series. This property is true of many other classes of organic compounds.

Returning to methyl alcohol — or, as it is now being called,

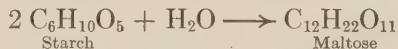
methanol — this alcohol is prepared commercially by the destructive distillation of hard wood, such as maple, beech, birch, etc. When wood is heated in retorts to about $400^{\circ}\text{C}.$, we obtain gases, an aqueous distillate known as “pyroligneous acid,” wood tar, and charcoal. The wood tar is a source of antiseptics (such as wood creosote), disinfectants, fuel, etc. The pyroligneous acid contains the methanol as well as acetic acid and acetone.

By the use of a catalyst in the shape of zinc oxide, and at a temperature of about $400^{\circ}\text{C}.$ and a pressure of about 150 atmospheres, methanol is being manufactured from carbon monoxide and hydrogen, the constituents of water gas:

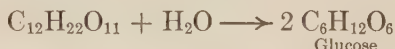


Methanol is a poisonous liquid and gives rise to blindness when taken internally. It is very largely used as a solvent in the manufacture of varnishes and lacquers, especially those containing shellac; as one of the solvents with ether for moving-picture films; in the preparation of denatured alcohol and various dyes; etc.

443. Ethyl alcohol, also called *alcohol*, *grain alcohol*, *wine spirit*, $\text{C}_2\text{H}_5\text{OH}$, is made commercially by a process called fermentation. Some starch-containing substance, such as a cereal (corn, rye, barley, etc.) or potatoes, is allowed to sprout; this develops within the cell a substance known as diastase, a typical catalyst, and called an *enzyme* because it is produced as a result of the activity of the cell. The diastase converts the starch into maltose:



Another enzyme, maltase, which is also present, converts the maltose into glucose:



Yeast is now added, and the enzyme zymase, present in the yeast, converts the glucose into alcohol and carbon dioxide:



The alcohol in the "wort," which is the name given to the liquor formed in the course of the fermentation process (carried out at from 25–30° C.), contains about 14 per cent alcohol, and is purified and concentrated by fractional distillation. Commercial ethyl alcohol contains about 95 per cent of alcohol. A still higher percentage of alcohol, *absolute* or nearly 100 per cent, may be obtained by the addition of calcium oxide or anhydrous copper sulphate (dehydrating agents) to the liquid and allowing it to stand a day or two; it is then distilled.

Grape juice, corn syrup, and molasses are already rich in glucose. Here the preliminary diastase treatment, consisting in the conversion of starch into glucose, is unnecessary.

Alcohol is extensively used in medicine in the form of "tinctures" (alcoholic solutions or extracts of medicinal substances). In the industries it finds a wide use: as a preservative, an antiseptic, in the preparation of denatured alcohol and various drugs and medicinals, as a solvent, in perfumery, in the manufacture of iodoform, chloroform, ether, etc.

444. Denatured alcohol is an alcohol which has been made unfit for drinking purposes (denatured) and external applications, but which can still be used in the industries. For "denaturing" purposes such substances as methanol, benzene, or acetone are added. Medicated alcohol is alcohol unfit for drinking purposes but suitable for external applications. Some of the substances which are mixed with alcohol for this purpose are tartar emetic, phenol, etc.

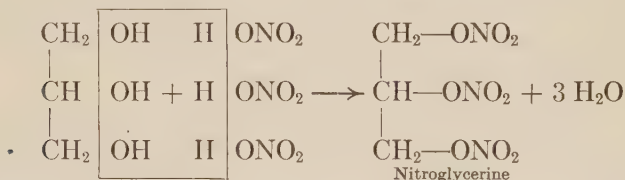
445. Glycerol, also called **glycerine**, $C_3H_5(OH)_3$, is an alcohol because it is a carbon compound containing hydroxyl groups. Its structure may be represented as



and it is produced as a by-product in the manufacture of soap.

Glycerol is a good solvent and dehydrating agent. It is used in medicine; as a sweetening agent; as a preservative for tobacco; in perfumery and cosmetics; in ink for rubber stamps; etc. When heated alone or in the presence of a dehydrating agent, acrolein is produced. The odor of burned fat is due to the formation of this substance.

When nitric acid is treated with glycerol under certain specified conditions, glyceryl trinitrate, or, as it is commonly called, *nitro-glycerine*, is obtained:

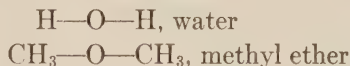


Concentrated sulphuric acid has to be added as a dehydrating agent to remove the water that is formed.

An alcohol very commonly met with in animal tissues (particularly in egg yolk and nervous tissue) is cholesterol, $\text{C}_{27}\text{H}_{45}\text{OH}$.

THE ETHERS

446. Ether. There is a large group of homologous compounds in organic chemistry known as the “ethers.” In structure they resemble the inorganic oxides. If water, H_2O , be regarded as hydrogen oxide, and if for each hydrogen atom we replace it with, say, the organic group CH_3 (the *methyl* group), we get

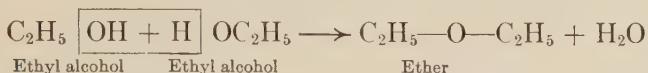


If instead of the CH_3 group we have the C_2H_5 group (the *ethyl* group), then we get



This ethyl ether is commonly spoken of as just “ether,” just as ethyl alcohol is very often referred to as “alcohol.” Ether is

manufactured by the reaction of sulphuric acid and ethyl alcohol. The equation may be expressed as



the sulphuric acid taking out the elements forming water.

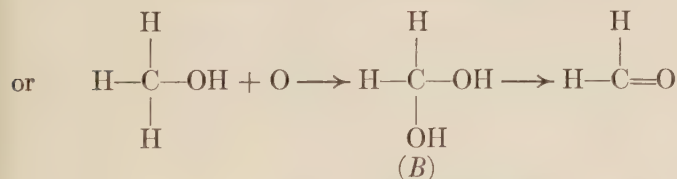
Ether is a highly inflammable liquid, being explosive when mixed with air and ignited. It is an excellent solvent for fats, oils, resins, alkaloids, etc., and as a solvent it is used very extensively. It is used with alcohol in the manufacture of guncotton. Due to its rapid evaporation it can be used for refrigeration purposes. As an anæsthetic it was introduced in surgery by Dr. Morton, a Boston dentist, in 1846.

THE ALDEHYDES

447. Formaldehyde, HCHO , is the simplest of a class of compounds known as “aldehydes.” The characteristic group of an

aldehyde is —CHO or $\text{—C} \begin{array}{l} \nearrow \text{O} \\ \searrow \text{H} \end{array}$, and one way of preparing an

aldehyde is by the oxidation of the corresponding alcohol. For example, when methyl alcohol is oxidized, we get



To explain the reaction, we suppose that at first an oxygen atom is added to the methyl alcohol molecule, giving the compound marked (B), which has, as you may see, two hydroxyl groups attached to the same carbon atom. Extensive experience has

shown that such a compound is very unstable, breaking up as indicated, with the elimination of water.

In actual practice, the methyl alcohol vapor and air are passed over copper gauze (catalytic agent?), the formaldehyde gas so obtained being dissolved in water making the *formalin* of commerce.

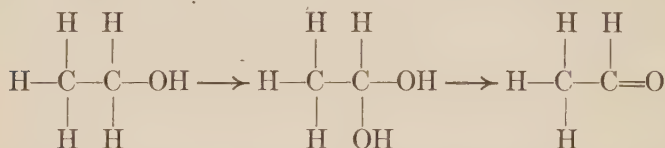
Formaldehyde is used as a disinfectant, food preservative, and germicide. It is also used in the manufacture of dyes (such as indigo), photographic films, and synthetic resins such as bakelite. It is supposed that in the formation of sugars in plants, formaldehyde is an intermediate product, thus:



and that the formaldehyde immediately polymerizes into a simple sugar:



448. Acetaldehyde, CH_3CHO , may be obtained by the oxidation of ethyl alcohol:



It is manufactured by passing acetylene into dilute sulphuric acid, using mercury salts as the catalyst. With chlorine gas it forms trichloroacetaldehyde, or chloral:

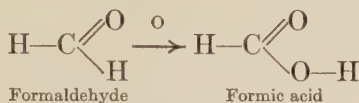


which is used as a drug to induce sleep, and, in larger doses, is also used as an anæsthetic.

THE ORGANIC ACIDS

449. Formic acid, HCOOH , is the simplest of a group of organic acids whose characteristic structure is $\text{—}\overset{\text{O}}{\underset{\text{O—H}}{\text{C}}}$. These acids

may be obtained by the oxidation of the corresponding aldehydes; for example:



It is manufactured by heating together sodium hydroxide and carbon monoxide, and is used in the manufacture of dyestuffs, in dyeing and finishing textiles, etc. The stinging by bees is due to the penetration of formic acid under the skin.

450. Acetic acid, CH_3COOH , occurs most commonly in *vinegar*, which, in turn, is produced by the fermentation of fruit juices. The sugar (in the juice) is first converted to alcohol (cider) by the enzyme *zymase* which is present:

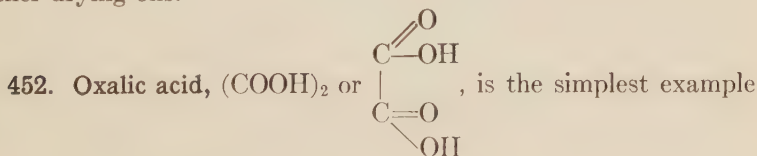


and the alcohol is next oxidized to acetic acid (vinegar) by oxidizing bacteria, present in the juice:



Vinegar contains from 3 to 6 per cent of the acid. Stronger solutions are obtained from pyroligneous acid. Apart from household uses acetic acid is also used in the manufacture of white lead, dyes, and drugs.

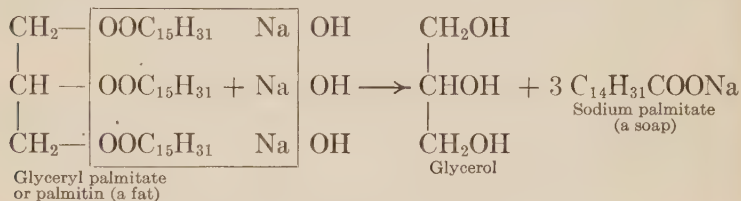
451. Butyric acid, $\text{C}_3\text{H}_7\text{COOH}$, found in rancid butter, and several other evil-smelling acids, are used in the preparation of some of the most delicate perfumes on the market. The fats are made up of combinations of glycerol and *palmitic acid* ($\text{C}_{15}\text{H}_{31}\text{COOH}$), *stearic acid* ($\text{C}_{17}\text{H}_{35}\text{COOH}$), and *oleic acid* ($\text{C}_{17}\text{H}_{33}\text{COOH}$). *Linoleic acid*, $\text{C}_{17}\text{H}_{31}\text{COOH}$, is found (in the form of a salt) in linseed and other drying oils.



of an acid containing two COOH (carboxyl) groups. It is used in the manufacture of dyes, in bleaching, in metal polishes, in tanning, etc. It is a highly poisonous substance.

SOAPS

453. Soaps. We have already spoken of palmitic, stearic, and oleic acids. The sodium or potassium salts of these acids are the common soaps with which we are so familiar. They are prepared by boiling fats or oils (which are combinations of glycerol with one or more of the acids just named) with sodium hydroxide (*saponification*):



The soap is separated from the glycerol by a "salting out" process, which means adding salt until the soap comes to the surface, when it is ladled off.

The solid soaps are sodium salts, whereas the liquid soaps are potassium salts.

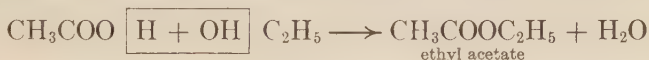
In the preparation of soap suitable for use with sea water, coconut oil is often used. Soaps made from "red oil," containing crude oleic acid, make excellent laundry products. Soap powders are rich in free alkali, containing, as a rule, not more than ten per cent soap and a certain quantity of inert material called a filler. In the preparation of scented soaps, perfume, color, and other substances are added. Tar soap contains tar, and antiseptic soap contains carbolic acid. The "grit" soap used to remove grease and tar from the hands contains pumice or silica and sometimes a little naphtha. Shaving soaps often contain an excess of oil or grease and sometimes glycerine, and a mild antiseptic

like borax. Zinc stearate is often applied to a chafed skin. Tooth powders or tooth pastes consist of soap mixed with chalk in a fine state of division (precipitated chalk), sugar, and a mild antiseptic.

454. Esters. An ester is either an inorganic or organic acid in which the acid hydrogen has been replaced by an R group, remembering that R stands for any organic radical, CH_3 , C_2H_5 , etc. They are in organic chemistry what the salts are in inorganic chemistry. A number of these esters are of great practical importance. For example, *glyceryl trinitrate*, erroneously called *nitroglycerine*, is an ester. It is prepared by the action of nitric acid on glycerol (sulphuric acid being used as a dehydrating agent):



and is the active constituent of dynamite. *Ethyl acetate* is an ester and is made by the reaction of acetic acid on ethyl alcohol:



It is used in the manufacture of artificial leather, as a solvent for nitro-cellulose, and in the preparation of photographic films.

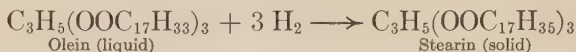
Many of the esters are used in the preparation of artificial fruit essences, flavors, and perfumes.

455. Fats and vegetable oils are esters; they are the combination of glycerol with organic acids (usually of high molecular weight). The glyceryl esters of stearic acid [stearin, $\text{C}_3\text{H}_5(\text{OOC}_{17}\text{H}_{35})_3$], palmitic acid [palmitin, $\text{C}_3\text{H}_5(\text{OOC}_{15}\text{H}_{31})_3$], and oleic acid [olein, $\text{C}_3\text{H}_5(\text{OOC}_{17}\text{H}_{33})_3$] constitute the main bulk of the fats and oils in food and of body fat.

Closely allied to fat in structure, but also containing the elements phosphorus and nitrogen, is *lecithin*, a substance found in animal tissues, particularly in the brain.

Oleomargarine is a mixture of tallow (the soft part of beef fat) and cottonseed oil, to which a small amount of butter and milk is often added.

456. Hydrogenation of oils. Liquid fatty oils can be hardened or solidified by treatment with hydrogen, in the presence of nickel as a catalyst; for example,

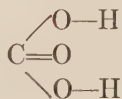


The olein is an unsaturated compound; therefore its molecule can take up hydrogen. These "hardened" fats (made from cottonseed oil, peanut oil, etc.) are now used extensively as lard substitutes, "Crisco," for example, and in the manufacture of soap.

457. Waxes, like fats, are esters, but in the place of glycerol, which the fats contain, the waxes have other alcohols (such as cetyl alcohol, $\text{C}_{16}\text{H}_{33}\text{OH}$). Carnauba wax is derived from a species of palm and is used in candle making and for adulterating beeswax. Lanolin (wool wax), obtained from wool grease, is used in pharmacy as a basis for ointments. Beeswax, obtained from the honeycomb of bees, and spermaceti, found in the head of the sperm whale, are used in candlemaking, shoe polish, etc.

SOME DERIVATIVES OF ORGANIC ACIDS

458. Phosgene. Carbonic acid, H_2CO_3 , may be written structurally thus:

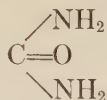


If the OH groups are replaced by Cl groups, we get phosgene:

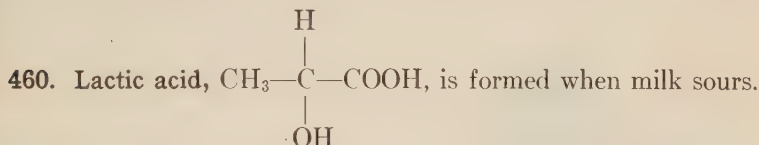


which, in practice, is prepared by passing chlorine and carbon monoxide over charcoal. Since it is very poisonous and can be easily prepared and liquefied, phosgene was extensively employed in the late war.

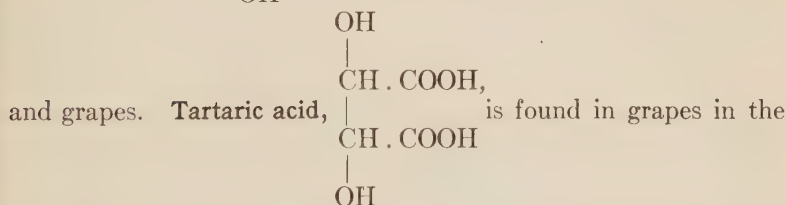
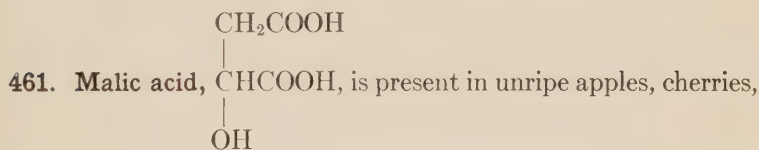
459. Urea may also be regarded as derived from carbonic acid in which each OH (hydroxy) group is replaced by an NH₂ (amino) group :



This compound is found in abundance in the urine of animals, and its amount is often an important indication of how much protein the body has utilized.



This is due to the action of certain bacteria present in the air on the milk sugar (lactose) present in milk. The "souring" of the milk is due to the lactic acid which has been formed.



form of potassium acid tartrate (cream of tartar), $\text{CH}(\text{OH})\text{COOK}$, which is also a constituent of some baking powders. Sodium potassium tartrate (*Rochelle salt*) is $\text{CH}(\text{OH})\text{COONa}$ and is used as

a laxative (as in "Seidlitz" powders). Potassium antimonyl

tartrate (*tartar emetic*), $\begin{array}{c} \text{CH(OH)COOK} \\ | \\ \text{CH(OH)COO(SbO)} \end{array}$ is used in medicine as

an emetic (a drug to produce vomiting). **Citric acid**, $\begin{array}{c} \text{CH}_2 \cdot \text{COOH} \\ | \\ \text{C(OH)} \cdot \text{COOH} \\ | \\ \text{CH}_2 \cdot \text{COOH} \end{array}$

is found in lemons, berries, etc. It is now often prepared on a large scale by the fermentation of glucose with certain fungi, and is used in lemonades and other beverages.

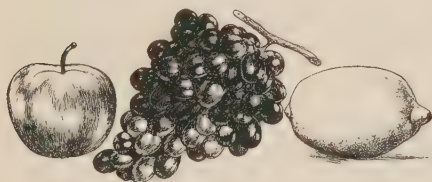
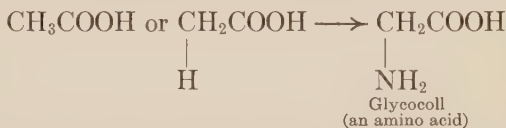


FIG. 279. Fruits contain organic acids.

AMINO ACIDS AND PROTEINS

462. An amino acid is a derivative of an organic acid, in which one of the hydrogen atoms in R is replaced by the amino (NH_2) group. For example, if in the place of R we write CH_3 , then RCOOH becomes CH_3COOH (acetic acid). Now



These amino acids are of great physiological importance, since it has been shown that all proteins are first broken down into amino-acids in the digestive tract before they are absorbed by the blood and utilized by the tissues. Some 16 different amino acids are known; but as their chemical structures are rather complicated, no attempt will be made to describe them here. However, it is of interest to note that Fischer has succeeded in recombining these amino acids with one another to form a substance resembling a protein — a great feat in organic chemistry.

463. Proteins. Proteins are now considered as combinations of amino acids. Despite the many proteins found in nature, they are all considered as derived from suitable combinations of the 16 amino acids. Some common proteins are the albumen of egg white, milk, and blood serum; the globulin of egg white and blood; the glutenin of wheat; the zein of corn; the keratin of hair; myosin in muscle; the mucin of saliva; the casein of milk; and the hemoglobin of blood.

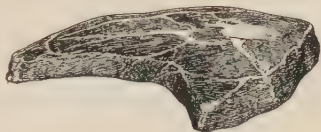
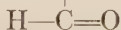


FIG. 280. Protein — lean meat.

CARBOHYDRATES

464. The common carbohydrates may be classified as monosaccharides, $C_6H_{12}O_6$ (glucose, galactose, fructose, etc.); disaccharides, $C_{12}H_{22}O_{11}$ (sucrose, maltose, lactose); polysaccharides, $(C_6H_{10}O_5)_n$ (starch, cellulose, glycogen, etc.). Chemically, the

simple sugars have the structure
$$\begin{array}{c} R \\ | \\ H-C-OH \end{array} \quad (A)$$
 and you will notice that they contain the hydroxyl (OH) and aldehyde (CHO) groups.



There is still another group, but it is not necessary in our discussion here.

465. Glucose, $C_6H_{12}O_6$, also called **dextrose** or **grape sugar**, with a structural formula resembling (A), is found in many fruits, such as the grape. It is a normal constituent of the blood and in pathological conditions (as in diabetes) accumulates to an abnormal degree in the blood and urine. Commercially it is prepared by treating starch with dilute acids (hydrolysis):



FIG. 281. Corn syrup contains glucose.

Glucose is used extensively in the making of confectionery, jellies, preserves, etc.

466. Fructose, levulose, or fruit sugar, $C_6H_{12}O_6$, has the same molecular formula as glucose, but its structural formula is slightly different. Fructose may be obtained by the hydrolysis of cane sugar.

467. Sucrose, cane sugar, $C_{12}H_{22}O_{11}$, is the ordinary table sugar of the household. On a large scale it is obtained either from the sugar cane or sugar beet. In either case, the juice is first separated from the insoluble fiber, then the albuminous material is precipitated and the acids are neutralized, and finally, after filtration it is decolorized with bone black and the filtrate is evaporated. When sucrose is hydrolyzed, a mixture of glucose and fructose, "invert sugar," is obtained :



and when heated above its melting point, caramel is formed. *Lactose* (milk sugar), $C_{12}H_{22}O_{11}$, occurs in milk. *Maltose* (malt sugar), $C_{12}H_{22}O_{11}$, is found in malt, which is the sprouted grain of barley.

468. Starch, $C_6H_{10}O_5$, is more correctly represented as $(C_6H_{10}O_5)_n$, where n represents a large number not yet determined. It is widely distributed in the vegetable kingdom and on hydrolysis or in the digestive tract is ultimately hydrolyzed to glucose :



With iodine starch gives a blue color (colloidal iodine).

469. Glycogen, animal starch, $(C_6H_{10}O_5)_n$, is found in large quantities in the liver, which is the storage house for the carbohydrates in the body.

470. Cellulose, $(C_6H_{10}O_5)_y$, is the chief constituent of the cell wall of plants. Cotton fiber is almost pure cellulose. Industrially, cellulose is of immense importance. It is the chief ingredient of cotton, linen, hemp, and paper.



FIG. 282.
Starch.

One common method of making paper is known as the sulphite process, which owes its origin to the pioneer work of Benjamin Tilghman, of Philadelphia. Here the wood pulp, such as spruce, is freed from bark, broken up into small pieces, and boiled with calcium sulphite. At this stage the resinous material is dissolved away, leaving the fiber in the form of pulp. From here on the details of the operation are mechanical rather than chemical. A "finish" to the paper is given by a process called sizing, which consists in applying resin or animal glue, thereby developing a smooth coating and making writing paper possible. Blotting paper and paper towels are not submitted to a sizing process. Glossy paper is obtained by the use of casein, the protein present in milk.



FIG. 283. Cotton is nearly pure cellulose.

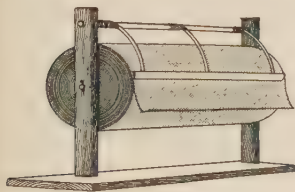
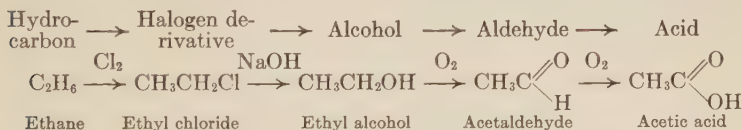


FIG. 284. Paper is chiefly cellulose.

To make parchment paper, paper is immersed in sulphuric acid of a very definite concentration for a certain length of time, and then thoroughly washed with water. The product has several times the strength of ordinary paper, due to the production of hydrocellulose, which forms a tough coating over the fibers; and is used to wrap butter and other food materials, to make "window envelopes," etc.

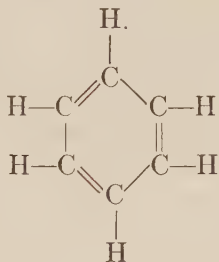
471. We can summarize some of the chemical transformations so far studied as follows:



THE BENZENE COMPOUNDS

472. So far we have discussed compounds which may, either directly or indirectly, be traced to methane, CH_4 . We now come

to a group of substances which include those out of which dyes and drugs are made, and which may be all traced back to a parent substance, *benzene*, C_6H_6 , which, after many years of careful study, is known to have this structural formula :



or  for short. Compounds derived from benzene are known as *aromatic*.

Benzene is obtained from coal tar and is used as a solvent for fats, resins, and in the manufacture of a large number of compounds.

473. Notice that benzene is a hydrocarbon, just like methane, and like the latter, gives rise to many homologues; for example, toluene, C_7H_8 ; xylene, C_8H_{10} ; mesitylene, C_9H_{12} ; etc.

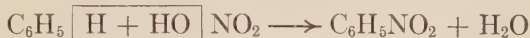
Benzene and its homologues form many compounds with the halogens, with nitric acid, with sulphuric acid, etc.; and these, in turn, are the starting point for many compounds of extreme industrial importance. Obviously, only a relatively few can be referred to here.

474. The simplest of the halogen compounds of benzene is chlorobenzene, C_6H_5Cl , which is made by combining benzene with chlorine in the presence of iron (which acts as a catalyst) and which is used for the manufacture of dye intermediates — substances which, in turn, are used to manufacture dyes. The simplest of the benzene compounds with sulphuric acid is benzene

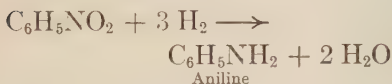
sulphonic acid, which is made by treating benzene with sulphuric acid :



The sulphonic acids are extensively used in the preparation of phenol (carbolic acid) and its derivatives. The simplest compound of benzene with nitric acid is nitrobenzene, which is made by treating benzene with nitric acid and adding sulphuric acid to remove the water formed (dehydration) :



It is largely used in the manufacture of *aniline* which is an important "dye intermediate." Aniline is prepared from nitrobenzene by reduction, using, as a rule, iron and hydrochloric acid for the purpose :



The first synthetic coal-tar dye, mauve, was made by Perkin some seventy years ago, using aniline as the starting material.

475. Acetanilide, formed by the action of acetic acid on aniline, is used in medicine to relieve pain and to reduce fever, and is a common constituent of headache powders. With carbon disulphide, aniline forms a compound (thiocarbanilide) which is used as an "accelerator" or "catalyzer" in the vulcanization of rubber.

476. But perhaps the most important reaction involving aniline is one known as *diazotization*, a process used so extensively in the

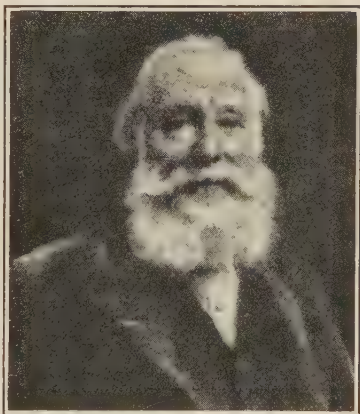
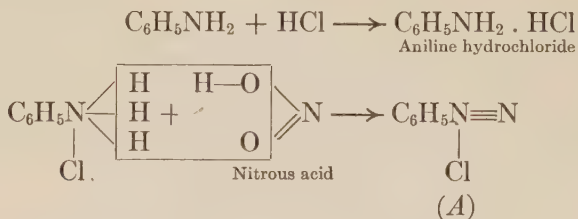


FIG. 285. Perkin, discoverer of synthetic coal-tar dyes.

industries in the manufacture of dyes. When aniline (or any homolog of aniline) is treated with nitrous acid at low temperature — the nitrous acid being prepared by adding hydrochloric acid and sodium nitrite to the mixture — we get benzene diazonium chloride (*A*):



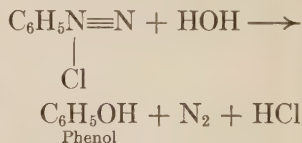
(*A*) and other diazonium compounds form the starting point for the preparation of hundreds and hundreds of different compounds.



Bakelite Corp.

FIG. 286. Some articles made of bakelite.

477. Phenols. When (*A*) is heated, *phenol* or *carbolic acid* is obtained:



which, in some respects, resembles an alcohol in structure, and which is the mother substance of many other phenols, just as methyl alcohol is the mother substance of many other alcohols. Much of the phenol of commerce, however, is obtained from coal tar. It is a powerful disinfectant and is also used in the manufacture of explosives, dyes, developers, resins, etc. One of the best known of these resins is **bakelite**, which is prepared from phenol

and formaldehyde, $\text{H} \cdot \text{CHO}$. After subjecting the bakelite to heat treatment it becomes insoluble, infusible, very hard, and very resistant, and has found extensive use in varnishes, enamels, lacquers, cements, pipe stems, cigar holders, handles, etc.

478. “**Tri-cresol**” is a mixture of the three cresols which are homologs of phenol, with the general formula $\text{C}_7\text{H}_7\text{OH}$. The “tri-cresol” has greater germicidal power than phenol and is less poisonous. *Lysol* is a preparation containing the cresols. *Thymol*, $\text{C}_{10}\text{H}_{13}\text{OH}$, another homolog of phenol, occurs in the oil of thyme and oil of mint, and is an important antiseptic. It has been extensively used in the treatment of hookworm. *Quinol* (or *hydroquinone*), $\text{C}_6\text{H}_4(\text{OH})_2$, which is used as a photographic developer, and *pyrogallol* (or *pyrogalllic acid*), $\text{C}_6\text{H}_3(\text{OH})_3$, used in estimating oxygen in gas mixtures such as air, are other well-known phenols.

AROMATIC ACIDS

479. The simplest of these acids is *benzoic acid*, $\text{C}_6\text{H}_5\text{COOH}$, which occurs in a number of resins and gums, and which, as might be expected, may be obtained by the oxidation of the corresponding aldehyde, benzaldehyde, $\text{C}_6\text{H}_5\text{CHO}$. This acid is used in the manufacture of dyes. Its sodium salt, sodium benzoate, is used as a food preservative. Phthalic acid, $\text{C}_6\text{H}_4(\text{COOH})_2$, combines with phenol to form *phenolphthalein*, an indicator almost as well known as litmus.

SEVERAL OTHER DERIVATIVES OF BENZENE

480. **Picric acid**, $\text{C}_6\text{H}_2(\text{NO}_2)_3\text{OH}$, is, as its formula indicates, a nitro compound and a phenol. It is manufactured from phenol in much the same way that nitrobenzene is manufactured from benzene — by the action of nitric and sulphuric acids. Like many other nitro compounds, picric acid is a powerful explosive.

481. **Chloramine-T** (or *chlorazene*) and **dichloramine-T**, are halogen derivatives of toluene, C_7H_8 , and were very successfully in-

troduced by Carrel and Dakin during the World War as active germicides and good irrigating fluids in the treatment of wounds.

482. Para-aminophenol, $C_6H_4(OH)NH_2$, has the structure and the properties of aniline and phenol and is used as a photographic developer and in the manufacture of dye intermediates. Phenacetin is derived from para-aminophenol and is used to reduce fever (antipyretic) and to relieve pain (analgesic).



Underwood and Underwood

FIG. 287. Remsen synthesized saccharin.

483. Vanillin, $C_6H_3 \cdot CHO \cdot OCH_3 \cdot OH$, is obviously both an aldehyde and a phenol and is used in perfumery and as a flavoring agent. It is present in vanilla.

484. Salicylic acid, $C_6H_4 \cdot COOH \cdot OH$, is both an acid and a phenol and is prepared from carbolic acid. One of its derivatives, *aspirin* (or acetyl salicylic acid), is very extensively

used in medicine; another, *methyl salicylate*, $C_6H_4 \cdot OH \cdot COOCH_3$, is the chief constituent of oil of wintergreen.

485. Saccharin, $\text{C}_6\text{H}_4(\text{SO}_2)(\text{CO})\text{NH}_2$, is prepared from toluene in a

number of successive steps. It is about 400 times as sweet as sugar and is, in certain cases, used as a substitute for sugar. It is extensively used as a sweetening agent in mouth-washes, toothpastes, etc.

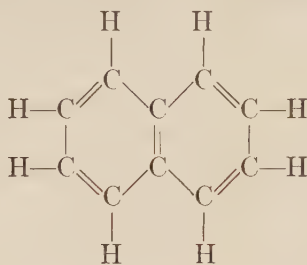
486. Gallic acid, $C_6H_2(OH)_3 \cdot COOH$, is similar to salicylic acid in being both a phenol and an acid. It is found in a number of plants and is also obtained from tannins by hydrolysis with acids. It is used in photography and in the making of ink. *Tannic acids*, found in gall nuts and other plants, are closely allied, chemically,

to gallic acid, and are largely employed in the making of leather, as mordants in dyeing and in the making of inks.

NAPHTHALENE, ANTHRACENE, AND THEIR DERIVATIVES

487. Our discussion of the aromatic compounds, so far, has centered around benzene and its derivatives. Now just as methane and benzene can give rise to many compounds, so can two other hydrocarbons, naphthalene and anthracene; and these, with their derivatives, we shall now discuss very briefly.

488. **Naphthalene**, $C_{10}H_8$, has the structural formula:



M. G.

from which it can be seen that the compound really represents a condensation of two benzene rings. It is obtained from coal tar and is used in the preparation of other naphthalene compounds, in moth balls, as an insecticide, and in the preparation of dye intermediates.

Like benzene, naphthalene can be chlorinated, sulphonated, nitrated, the nitro compounds can be reduced to form amino (NH_2) compounds of the type of aniline, the amino compounds can be diazotized, the diazonium compounds can be heated and the corresponding phenols (or naphthols, as they are here called) obtained, etc. In fact, for every compound of benzene there is an analogous compound of naphthalene.

489. **Anthracene**, $C_{14}H_{10}$, is a condensation of three benzene nuclei:



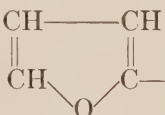
and is also obtained from coal tar. It is largely used for the preparation of *anthraquinone*, which is obtained by oxidizing anthracene. Anthraquinone is used in the manufacture of dyes (such as alizarin) and dye intermediates.

TERPENES

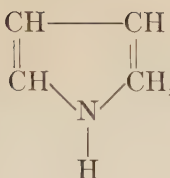
490. The terpenes are complex hydrocarbons, derived from a mother substance with the formula $C_{10}H_{16}$, which are obtained from such substances as camphor, oil of turpentine, and "essential oils" (products obtained when certain plants are distilled with steam). They are used extensively as flavors, in perfumery, and in medicine. They are closely related to cymene, $C_6H_4 \cdot CH_3 \cdot C_3H_7$, which is a benzene derivative. Examples of terpenes are: pinene, the chief constituent of oil of turpentine; menthol, which occurs in oil of peppermint; and camphor, which is obtained from the camphor tree and which is largely used in the manufacture of celluloid and in pharmaceutical preparations.

HETEROCYCLIC COMPOUNDS

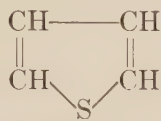
491. So far we have considered "ring" compounds (benzene, naphthalene, anthracene, and their derivatives) which have the same element in the ring, namely, carbon. We shall now take up, very briefly, a number of "ring" or "cyclic" compounds in which elements other than carbon are present. Such compounds are known as *heterocyclic*.

492. Furfural,  as may be seen, is an

aldehyde, and shows all the characteristic properties of one. Commercially it is prepared from corncobs and is used in the manufacture of synthetic resins (similar to bakelite), disinfectants, deodorizers, etc.

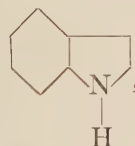
493. Pyrrole, , is present in coal tar. The pyrrole

ring is present in a number of alkaloids.

494. Thiophene, , is present in crude benzene and is the mother substance of almost as many compounds as is benzene itself.

495. Pyridine, , is found in coal tar and is used to denature alcohol.

496. Quinoline, , is a condensation of one benzene and one pyridine ring and is present in coal tar. The quinoline ring is present in a number of alkaloids.

497. Indole, , repre-

sents a condensation of the benzene and pyrrole rings and is a product formed in the intestines during the putrefaction of proteins.

498. Indigo, an important dye obtained from the indigo plant, may be regarded as derived from indole. It is now very largely prepared synthetically in the laboratory, the starting point being either naphthalene or aniline.

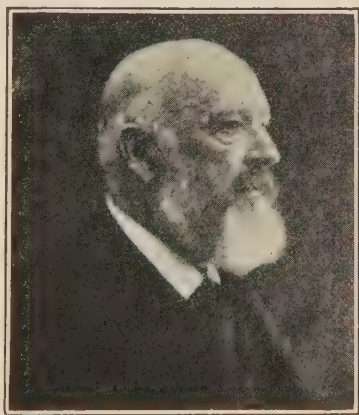


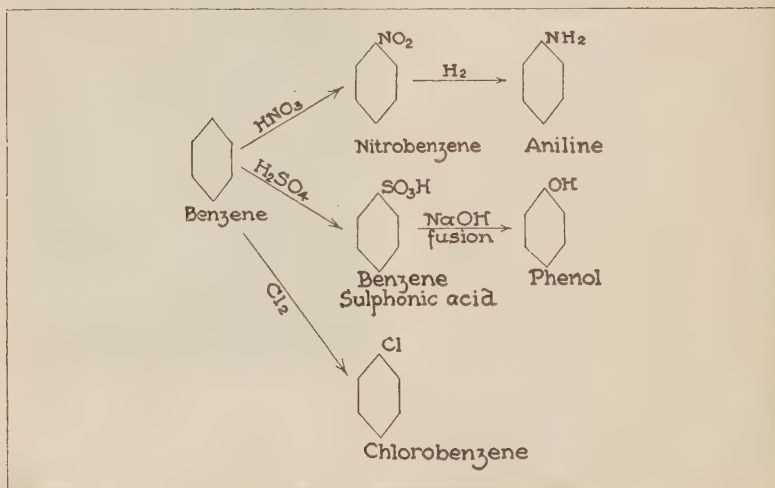
FIG. 288. Baeyer, who synthesized indigo.

VEGETABLE ALKALOIDS

499. These are complex nitrogen-containing substances occurring in plants which have powerful effects on plants and animals. They often contain pyrrole, pyridine, and quinoline groups in their structure. Some of the commoner ones are *nicotine*, present in tobacco leaves; *atropine*, used for dilating the pupil of the eye; *cocaine*, sometimes used as a local anæsthetic; *quinine*, used in the treatment of malaria; *strychnine*, sometimes used in small doses as a tonic; and *morphine*, the chief alkaloid of opium. At present there are on the market more than 1500 distinct dyes and drugs, all of them the product of the labors of the organic chemist.

Dyes and drugs, as you may notice in this chapter, seem to be the chosen field of the scientist who specializes in organic chemistry; just as the preparation of such substances as "synthetic" ammonia, sulphuric acid, and "synthetic" gasoline seem to be the chosen field of another kind of chemist — the *physical* chemist.

500. Some of the reactions just studied in the chemistry of the aromatic compounds (benzene and its derivatives) can now be summarized:



SUPPLEMENTARY READING

- Perkin Medal Award to Dr. W. M. Burton (for work on the chemistry of petroleum), *Journal of Industrial and Engineering Chemistry*, vol. XIV (1922), p. 159; also same journal, vol. X (1918), p. 484.
- D. L. Davoll, "The Manufacture of Sugar," *Chemical Age*, vol. XXXII (1924), p. 11.
- "Dye Symposium," *Industrial and Engineering Chemistry*, vol. XVI (1924), p. 409.
- C. O. Johns, "Petroleum," *Industrial and Engineering Chemistry*, vol. XV (1923), p. 446.
- G. J. Esselen, "Cellulose Chemistry," *Industrial and Engineering Chemistry*, vol. XVIII (1926), p. 1031.
- L. C. Newell, "Faraday's Discovery of Benzene," *Journal of Chemical Education*, vol. III (1926), p. 1248.
- William Foster, "From Atom to Life," *Journal of Chemical Education*, vol. III (1926), p. 1391.
- F. A. Howard, "Fifty Years in the Petroleum Industry," *Industrial and Engineering Chemistry*, vol. XVIII (1926), p. 936.
- H. E. Howe and F. M. Turner, *Chemistry and the Home* (Charles Scribner's Sons, 1927), p. 131 (silk), p. 154 (paper), p. 190 (rubber).
- M. G. Luft, "Rayon," *Chemistry in Industry*, vol. II, p. 305.
- M. H. Ittner, "Soap," *Chemistry in Industry*, vol. II, p. 360.

QUESTIONS

1. Explain the derivation of the word "organic."
2. Why should the chemistry of the carbon compounds be a separate branch of chemistry?
3. Write the structural formulas for propane and butane.
4. Consult some one reference book and prepare a short account of the history of petroleum.
5. Write the structural formulas for ethylene and acetylene and give one method of preparing each of them.
6. To what extent has synthetic rubber become a reality?
7. Write the structural formulas for chloroform, iodoform, and carbon tetrachloride.
8. What does the organic chemist mean when he speaks of an "alcohol"? Select some one alcohol, write its structural formula, show how to prepare it, and give some of its important uses.
9. What is the characteristic group of (a) an ether, (b) an aldehyde,

(c) an acid? Give an example of each one of these and cite one characteristic property in each case.

10. How would you prepare soap?

11. What is an ester? Name two esters, give their formulas and their uses.

12. Explain the phrase "hydrogenation of oils."

13. What is the difference between a wax and a fat? Give an example of each.

14. Give the structural formulas for (a) phosgene, (b) urea, (c) lactic acid, (d) malic acid, (e) tartaric acid, (f) citric acid. Give a use for each one of these substances.

15. What is an amino acid and what relation has this acid to a protein?

16. Classify the carbohydrates into three main divisions and give the general formula for each division.

17. Give the structure of a simple sugar and point out the presence of two important groups in the molecule.

18. Give one other name for each of the following and one use for each substance: glucose, fructose, sucrose, lactose, maltose, glycogen.

19. What products are obtained upon the hydrolysis of (a) sucrose, (b) starch?

20. What is "rayon" and how may it be obtained?

21. Why are the cellulose nitrates of such enormous industrial importance?

22. Give the structural formula for benzene.

23. Name the compounds formed when benzene reacts with (a) chlorine, (b) nitric acid, (c) sulphuric acid; when nitrobenzene is acted upon by a reducing agent.

24. What is formed when aniline is treated with nitrous acid at low temperature?

25. What is a phenol? Name several phenols and give one important use for each.

26. Name and give the formula of an aromatic acid.

27. Write the formula for sodium benzoate.

28. Name the characteristic groups present in (a) picric acid, (b) para-aminophenol, (c) vanillin, (d) salicylic acid, (e) gallic acid. Give one use for each of these substances.

29. Name two hydrocarbons other than methane and benzene which give rise to a large and important group of organic compounds.

30. What is a terpene? Give two examples.

31. What characteristic element within the molecule do we find in (a) pyrrole, (b) thiophene?

32. How might you prove that furfural is an aldehyde?

33. Name three alkaloids and give one use for each.

CHAPTER XXXI

THE CHEMISTRY OF FOOD

501. The question of what to eat involves much knowledge. You must know the composition of the foodstuffs; you have to be familiar with the organs that digest them, and with the chemistry and physiology involved; you have to know the fate of these foodstuffs in the body, and how they affect the different organs of the body. Bitter experience, coupled with a peculiar something which for want of a better name we call "instinct," has been our guide in the past; but since the days of Lavoisier we have gradually built up a science of eating, or, as it is commonly called, a science of nutrition. Here we can only discuss the more simple chemical factors involved.

502. Lavoisier. Lavoisier showed that substances were oxidized in the body by combining with oxygen and giving out carbon dioxide, in much the same way that a piece of wood or coal is burnt. With the help of his scales he introduced quantitative measurements. As the idea of *quantity* underlies all the sciences, we may justly speak of Lavoisier as the founder of the science of nutrition. He was also the founder of the modern science of chemistry, as you may recall.

503. The calorie. When substances burn, they give off heat. How are we to measure this heat? We do so by making use of the *calorie*, which may be defined as that amount of heat necessary to raise the temperature of one gram (about $\frac{1}{36}$ of an ounce) of water one degree centigrade. The large calorie is 1000 times this number. Whenever we deal with food values, we use the large calorie; so that in this chapter, whenever the word "calorie" is

mentioned, it must be understood as meaning the amount of heat necessary to raise the temperature of one kilogram (1000 grams) of water one degree centigrade.

Lavoisier has taught us that foods are oxidized in the body. This means that, among other things, heat is evolved. Such heat, in turn, can be measured in terms of calories.

The average of thousands of experiments in different parts of the world leads to a figure closely approximating 2700 as the number of calories per day needed by man. This is merely another way of saying that the average man oxidizes enough food each day to yield 2700 calories; and therefore the body needs to be replenished by this amount.

Dr. Holt, the famous children's specialist, finds from his experience that a child one year old needs 100 calories for every kilogram of its body weight (the kilogram being $2\frac{1}{2}$ pounds); that the needs slowly decrease to 80 calories until the tenth year is reached, and that, after the growth cycle is completed, 44 calories per kilogram suffices. For example, a boy 15 years old, weighing, say, 55 kilograms (about 121 pounds), may need 65 calories per kilogram weight; or a total of $65 \times 55 = 3575$ calories.

504. The carbohydrates. These are the primary energy producers. The sugars (cane sugar, candy, jellies), starch (in grains, bread), and cellulose (in potatoes, fruits, and vegetables) are common examples of carbohydrates. Various experiments have shown that though the fats and the proteins (which we are about to discuss) also serve as energy or heat producers, the complete absence of carbohydrates from the diet gives rise to various nutritional disturbances (an "upset stomach," for example). It should also be remembered that the laxative properties of foods are largely due to the carbohydrates (cellulose) in them; and man when in normal health needs no other laxative than the food he eats.

505. Fats. Butter is perhaps the commonest variety of fat. Various oils, lard, portions of meat, and cream are other examples. It has already been pointed out that fats, like carbohydrates, are

energy producers; but the fats act as the reserve forces. The carbohydrates are the active fuel producers and the first in the field; the fats are called upon when the carbohydrate supply is well-nigh exhausted. The question whether fats are really a necessary part of the diet has not yet been definitely answered. Experiments on animals would tend to show that the amount needed, if anything, is very small. On the other hand, the Germans, during the late war, attributed many of the consequences of their poor health to the lack of fats. Fats in foods, in addition, invariably include a group of substances known as "lipoids," which, from their great abundance in brain tissue and their presence in all cells, make it certain that they are of the utmost importance to our welfare; and most fats contain, also, appreciable quantities of one of the vitamins.

506. Protein. We cannot do without protein (present in large percentages in meat, fish, milk, and eggs, and in smaller percentages in cereals and vegetables); so much is certain. Protein is the only one among our foodstuffs that contains nitrogen, and our tissues need nitrogen — not, unfortunately, in the free state, for then the air could supply us with all that we need; but in the combined and more expensive form of protein.

Having decided what our total calorific needs are (probably in the neighborhood of 2700 calories for adults, and several hundred calories more for active and growing children), we must next decide how to apportion our foodstuffs; how much protein and how much fat and carbohydrate are we to take. Roughly speaking, we may say that the proportions are 2 parts of protein to one part of fat to 10 parts of carbohydrate. But if we are to be guided by the experience of mankind, as well as by a number of recent experiments performed by scientists, we would suggest the use of ten per cent of our total food in the form of protein; the rest may be distributed among the fat and carbohydrate, with a decided preponderance in favor of the latter (perhaps one part of fat to ten parts of carbohydrate).

507. Mineral salts. These are the inorganic constituents of foods popularly known as "ash." Most of the elements included under this heading, such as potassium and magnesium, are apparently needed in such small quantities that a person fails to get the amounts needed only in acute starvation. The fats, proteins, carbohydrates, and water all contain small quantities of mineral matter, and only a total deprivation of food will stop the intake of these inorganic salts. But there are some inorganic constituents, like compounds containing calcium, phosphorus, and iron, that man needs in fairly large quantities, and a poorly selected or restricted diet may be deficient in these elements; just as a poor soil may be deficient in such elements as phosphorus, nitrogen, and potassium, which the plant needs.

508. Iron. Though all the cells of the body contain traces of iron, the largest amount is found in the hemoglobin (blood pigment) of the red blood cells. The amount of iron in the body is not large — about one tenth of an ounce. Not more than one two-thousandth of this amount is eliminated daily, so that the actual quantity of iron needed is small; but it must also not be forgotten that one of the richest iron-containing foods, egg yolk, contains but one part of iron for every 12,500 parts of yolk.

Milk is comparatively poor in iron. This seems rather strange at first sight, when we remember that milk is the sole source of nourishment of the infant. Light has been thrown on the subject by the discovery that, weight for weight, the infant's tissues are much richer in this element than are those of the adult. The suckling animal stores iron, which is obtained from its mother even before birth, and makes use of this reserve supply during the first few months. This, of course, necessitates offering the mother food abundant in the element.

509. Calcium. "The ordinary mixed diet of Europeans and Americans," writes Professor Sherman, "at least among dwellers in cities and towns, is probably more often deficient in calcium than any other element." Two per cent of the body weight is in the

form of calcium, and more than 99 per cent of the total amount is found in bones. In the structure of bones, calcium and phosphorus fight for supremacy: both are essential; when either is present in insufficient amount, or when either the one or the other is imperfectly assimilated, we often get the picture of a "rickety" child — a child whose bones are soft and flexible, and who is bow-legged. The most abundant source of calcium is milk and its products. A little over a pint of milk a day — a quart in the case of children because they are growing — is enough to supply all the calcium needed by the body. Vegetables, such as cabbage, turnips, and carrots, and fruits, such as the orange, are also rich in calcium. It is of importance to remember that bread and meat are particularly poor in this element; which is another reason for adopting a varied diet.

510. Phosphorus. Much of what has been said about calcium applies to phosphorus: both are essential elements, and both take an active part in bone and teeth formation. Cheese, eggs, beans, and peas are rich in phosphorus. The egg is rich both in iron and phosphorus.

511. Iodine. The importance of iodine as a constituent of the thyroid, a gland situated in the neck, has only lately been emphasized. The prevalence of goiter, a disease due to an insufficient supply of iodine, among children in the Great Lakes section of this country illustrates how the body may receive an insufficient supply of the element; and its administration (usually in the form of one of its compounds, such as sodium iodide) brings about a cure.

512. Vitamins. These are substances present in exceedingly small quantities in most of the foods we eat. They are essential to life. A diet which includes milk, vegetables, and fruit in fair abundance assures an adequate supply of these necessary substances. Several of these vitamins have been identified.

A clear conception of what vitamins are may be gathered from an experiment due to Professor Hopkins. He took two sets of rats — which we shall call A and B — in about the same stage of develop-

ment, and of about the same weight, and fed A with carefully purified casein (to represent protein), lard (to represent fat), and sugar (to represent carbohydrate), together with mineral salts and water. At the time of the experiment (some 15 years ago) this was believed to constitute an adequate diet. Group B was given

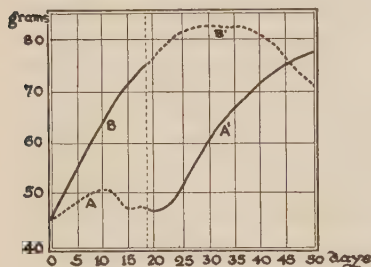


FIG. 289. Hopkins' experiment dealing with the discovery of vitamins.

the same diet, but each rat received, in addition, 2 cc. (roughly $\frac{1}{15}$ ounce) of milk per day. The rats belonging to series A lost weight and showed decided symptoms of illness; those belonging to series B gradually gained in weight. At the end of the eighteenth day of the experiment, the feedings were reversed:

this time group A received the milk and group B did not. From now on, group A gained in weight and thrived, whereas group B lost weight and showed symptoms of disease.

Now what conclusions are we to draw from such an experiment? Obviously that the animal needs something in addition to fat, protein, carbohydrate, mineral salts, and water; and that this "something" was supplied by adding 2 cc. of milk to the diet. This "something," which the chemist has not yet isolated, but which is necessary for life and is present in such small quantities, we call a *vitamin*.

Vitamin A is present in abundance in milk, butter, egg yolk, cod-liver oil, and, to a lesser extent, in beef fat and in many vegetable foods (lettuce, spinach, cabbage, carrots, potatoes, etc.).

Vitamin B is very abundant; nearly all the natural foods contain some of it. Yeast is particularly rich in this vitamin; so are milk and orange juice. The cereals contain it, but only the outer layers, so that in patent flour there is much less of this vitamin than in whole wheat flour.

Vitamin C is found in fruits and vegetables. The orange and the tomato are particularly good examples.

Very recent investigations tend to show that there are probably two other vitamins, which have been called, provisionally, vitamins D and E.

Diseases due to lack of vitamins. The absence of vitamin A in the diet gives rise to an eye disease known as xerophthalmia.

Such a disease was not uncommon among infants in Europe during the late war. The absence of vitamin B gives rise to a disease known as "beri-beri," in which the nerves are at-

tacked, paralysis developing sooner or later, followed, ultimately, by death. In experiments on pigeons, in which the disease can be induced by feeding them on a diet of "polished" rice, the subsequent addition to the diet of very small quantities of yeast (rich in vitamin B) will bring about remarkable cures within a comparatively short time. When vitamin C is absent, scurvy develops. In scurvy the gums become spongy and the teeth very loose. This disease at one time was very common among sailors who had to live for long periods on salted foods — on foods, that is, which were not fresh; and is common enough even in our own times in famine-stricken districts.

Rickets, in which the bones of the body are weak (the bow-legged youngster is often a good example), is by some regarded as being due to the absence of vitamin D, though there is no unanimity of opinion on the subject. Very recent work has shown that this disease, common among children, can be cured in a number of ways. One way is to expose the child to the action of the sun's rays. Another is to feed it with cod-liver oil.



FIG. 290. (1) Pigeon suffering from beri-beri (absence of vitamin B). (2) Same bird after treatment with yeast which contains vitamin B.

Scientists have succeeded in ascertaining the substance in cod-liver oil which is responsible for the cure. It is called "cholesterol" and is quite a common constituent of cellular material, particularly of brain cells. The most recent work leads to the belief that an impurity, present in very small quantities in ordinary cholesterol, is really the effective agent.



FIG. 291. Dog suffering from rickets.

Vitamin E, a fifth vitamin, has recently been identified (in wheat germ and other substances), and has been shown to play a very interesting rôle in reproduction.

513. Water and oxygen. Both water and oxygen are essential foods, for without either, life is impossible. The water is necessary for cell construction and for the fluids of the body. The oxygen makes combustion possible; so that the foods, or their products, coming in contact with this element, are oxidized, and so liberate the heat necessary to propel the human engine.

514. A wholesome diet. In planning our diet we ought, wherever possible, to plan it around milk as the nucleus. Milk contains the three classes of foodstuffs (fat, carbohydrate, and protein), mineral salts, and all the vitamins to a greater or less extent. A certain amount of vegetables and fruits should always be included — the particular kind depending upon the season. Aside from their vitamin content, thereby supplementing the supply in milk, fruits and vegetables contain appreciable quantities of mineral matter, and are helpful as laxatives; and vegetables also contain carbohydrates.

With milk as the nucleus of the diet, and moderate quantities of fruit and vegetables added to the milk, we may still be somewhat deficient in protein requirements. These may be obtained from meat, fish, eggs, or any one of the cereals (wheat, rye, barley, etc.). The cereals most commonly eaten, as in bread, are also rich

in carbohydrate, though poor in fat. Fat and vitamin A, if needed — and they will be needed particularly by the growing child — may well be obtained, in part, at least, from butter.

515. An important economic problem. One of the problems which forever occupies the attention of the chemist is to prevent waste. It has been pointed out in several portions of the book how he prevents waste in the industries. He has now turned his attention to preventing waste in food.

It has long been known, for example, that one way of preventing milk from spoiling is to evaporate part of the water (under reduced pressure) and to add sugar. Such “sweetened condensed milk” can be sealed in cans and kept for a very long time. This has been practiced for many years. Recently the chemist has devised a method whereby milk can actually be converted into a dry powder without losing any of its essential qualities. The powder, of course, keeps indefinitely. The principle of this process consists in first evaporating the milk under reduced pressure (as is done in the preparation of “condensed” milk), thereby getting rid of a large percentage of the water present, and then forcing the product under high pressure through an atomizer into a chamber in which the air has been heated. A white powder collects on the floor. This is essentially milk minus its water.

An even more extensive plan to preserve essential foodstuffs has been developed by chemists through a general process of *dehydration*. Fruit and vegetables can now be dried under carefully regulated conditions of temperature and moisture, so that not only can the dried product be kept indefinitely, but when ready for use and warm water is added, “the product takes on most of the characteristics of the fresh fruit or vegetable.”

516. Digestion of foods. Just what happens to the various foods we eat from the time they are taken into the mouth? That is a question which many people ask but which not many people can answer.

The human machine is often compared to a steam engine. The

analogy is a fairly good one, but it has its limitations, and these have to be understood. The food we eat, we are told, is like the coal with which we feed the engine; and just as the coal is the source of power in the engine, so is the food the source of power in the body. You will presently see that in many respects this analogy does not hold.

517. Digestive organs. In the first place, before our food can serve as the source of power, it has to undergo a series of preliminary changes in its chemical make-up; the chemist says we have to break these complex materials down into simpler units. This work is done in a tube-like arrangement known as the "gastro-intestinal canal." Understand this well: the mouth, the stomach, the intestines, etc., are primarily important because in them the food is so changed, so simplified, that it can then be taken up by the blood and carried to the various cells of the body. This immediately tells us why the *digestive organs*, such as the stomach and intestines, are so important for our well-being. If one or more of them stop functioning, then the preliminary simplification of the food cannot be carried out satisfactorily, the blood cannot take over food so ill-prepared, and the cells of the body do not get their proper nourishment.

518. Enzymes. The next question we must answer is what, in the gastro-intestinal tract, causes these changes in the food? How is the food broken down into units so simple that the blood takes them up with avidity and carries them to the cells — a thing the blood cannot do when it is offered food before it has undergone treatment in the gastro-intestinal tract? The reason for this is as follows: around the mouth and the stomach and the intestines there are factories known as "glands," whose function it is to manufacture products ("enzymes") which have the specific property of acting on foodstuffs and simplifying them. Some of these enzymes, such as those produced in the mouth, attack carbohydrates; others, as those in the stomach, attack proteins; still another attacks fats. There are no enzymes which act on mineral

salts; these substances are sufficiently simple for the blood to absorb them just as they are.

Let us make this enzymic action clearer by an analogy. A house has to be built. We require bricks and wood for the purpose. Bricks have first to be manufactured; the wood is first obtained from the trees. Then by intelligently combining these bricks and wood, the bricklayer and the carpenter build the house. The trees and the raw material out of which the bricks are manufactured represent the foods before they are eaten. The bricklayers and the carpenters are the "enzymes" — the workmen who prepare and arrange the materials so that they can be used to build the house (in the body, to build, and repair, body tissue). There are flaws in the analogy, but in its broad outlines it holds true.

519. We are now at the point where the food in the digestive tract has been so changed that it is in a condition to be absorbed by the blood and thereby carried to the various tissues of the body (for you must know that the blood is the great railway system of the body). The food has now been changed beyond recognition, and yet, up to the point of its absorption by the blood, *it has been of no actual use to the body*. From now on, however, the situation changes. The blood carries the altered food products first to the liver, where a number of substances of no use to the cells are retained and then discarded. The blood then brings the rest of the materials into the general circulation, in the course of which the various cells make use of their selective property of taking away from the blood those materials — the changed food products — which they need for their own cycle of activity.

520. Cells. We have referred a number of times to "cells." A cell is the simplest unit of living matter. The simplest of living creatures, such as the amœba, is a unicellular organism. Man, on the other hand, is of a much more complex nature, for he is made up of millions and millions of cells. In each one of these cells the sum of properties which we call "life" manifests itself; — materials are taken from the blood, reconstructed into living tissue, and

waste materials are given up to the blood. In the growing animals, the materials taken from the blood not only serve to replace waste tissue, but also to create new tissue.

At this point you can see that the analogy of cellular activity to a steam engine fails; for we have yet to discover the steam engine that can repair itself; we have yet to discover the steam engine that can be repaired by the use of coal or its products; and we have yet to discover the steam engine that can enlarge itself in size.

521. The cells, then, are the center of life in the body. Complex syntheses and analyses — the mechanism of which is little understood — take place within them: food is taken in, waste products are given off, and life continues for a space of time. Oxygen, the gas present in the air, and which we take in when we breathe, plays an irreplaceable part in these changes, for at one stage of these operations an "oxidation" goes on which is really related to the "burning" we watch in a fire; and just as such burning is impossible without oxygen, so oxidation in the body is impossible without oxygen. The analogy may be carried a step further; just as when coal is burned in a fire carbon dioxide and moisture are produced, so we can show that two of the waste products carried away from the cells and largely exhaled through the breath are also carbon dioxide and water. But in addition there are other non-gaseous waste products which the blood gives up to the kidney, and which are ultimately expelled from the body.

522. Summary. Summing up, then, we may say that the food which we eat has first to undergo a preliminary change in the digestive tract; the changed material is then absorbed by the blood, and carried to the various cells, which take what they want; these cells in turn building new tissue or replacing old, discarding the waste products also through the blood.

523. A problem. But there are obvious and enormous gaps in our knowledge. The chemist can analyze the food that goes into the body and he can also analyze the waste products that come out of it. He can, in other words, tell us something about the

start of the journey and something about its finish, but of the various intermediate steps in the process, those wonderful intermediate steps which go on in the cells themselves, neither he, nor his close co-worker, the physiologist, nor the recent issue of these two workers, the physiological chemist, can tell us much at present. And as to those delicate metabolic activities of the brain, man's supreme inheritance, we know very little.

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QUESTIONS

1. What was Lavoisier's contribution to the science of nutrition?
2. What is a calorie? How is the calorie used in measuring food requirements?
3. Give two common examples each of a carbohydrate, fat, and protein, and state what particular purpose such foods serve.
4. Name four inorganic elements which are important constituents of our diet and state what the function of each is. In each case name a food rich in the element in question.
5. What is a vitamin? How many vitamins have so far been identified?
6. Of what value are water and oxygen?
7. Describe briefly what you would consider a wholesome diet.
8. What important rôle do our digestive organs play?
9. What is an enzyme?
10. What food have you eaten to-day which contains vitamins A, B, and C?

SUPPLEMENTARY CHAPTERS

CHAPTER XXXII

PLASTICS

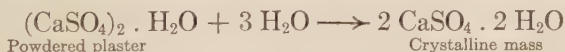
524. Plasticity is the ability to be molded and to retain the shape imparted, and many different substances show the property. First we must inquire whether there are any principles or rules governing the action of plastics, for this will enable us the better to understand the whole matter. The simplest plastic is the liquid, just as it is becoming a solid. For example, wax and many other substances, when heated, become plastic. Even metals, like lead and gold, if heated properly, can be worked easily. In fact, almost everything becomes soft and pliable when heated to the proper temperature. Why do most materials in changing from the liquid to the solid state (except perhaps "true" solutions) pass through an intermediate or plastic condition? Put in this form, the question is more readily answered. It is because the substance is colloidal at that time. The plastic, then, is a gel, and the same principles that govern gels control also plastics. How to prevent setting, to slow it up, to hasten it, are the same for plastics as for gels. Plasticity is increased by aging, tannin, and aluminum chloride; retarded by alkalis.

525. Clay. All clays show Brownian movement, and the plasticity of the particular sample depends only upon the amount of colloidal material present. It can, for example, be measured by the amount of malachite green (a positive colloid) adsorbed by the clay. Solution (or peptization) is increased by alkalis, and gelation by acids or salts. For instance, clay is made to flow, without adding too much water, by a trace of sodium hydroxide. This peptizes

the clay, and makes it less plastic, so it can be poured. For this reason, acids and salts are used to increase plasticity and to thicken suspensions of glazes and enamels. Most of you will recall how the Egyptians made bricks with straw (Figure 162). The straw contains tannic acid, which makes the clay more plastic and hastens the setting. When heated to 900° , complete coagulation sets in. Finally, crystals are formed and the mass is rigid.

To produce pure clay for white porcelain, advantage is taken of the fact that clay is negatively charged, while ferric hydroxide, rust, is positive. Electrophoresis causes the clay particles to migrate toward the positive electrode, while the ferric hydroxide collects at the negative electrode.

526. Cements. Many of the plastics are called cements, when they come into the market as powders. Many of them are made by heating crystalline substances and driving off the water of hydration. This leaves a powder, which, on absorbing water, proceeds to form crystals again; passing, of course, through the intermediate colloidal stage, where it is plastic. Plaster of Paris and hydraulic cement are examples of this type. The setting of plaster may be shown by equation:



Powdered plaster

Crystalline mass

Plaster is made from a natural crystalline calcium sulphate, called gypsum. On heating, water is driven off, and a white powder, plaster of Paris, results.

Hydraulic cement is made by heating together clay and limestone almost to fusion. This mass is then ground to powder:



Keystone View Co.

FIG. 292. Dentist using an amalgam to fill teeth.

our cement. Mixed with sand, water, and broken stone it makes *concrete* or artificial stone.

Many *amalgams* (mercury alloys) are used as cements. One part of tin rubbed with four parts of mercury makes a plastic mass which sets in a few days. Copper amalgam, in boiling water, is plastic, while when cold it is so hard it can be polished.



Keystons View Co.

FIG. 293. Native collecting rubber latex.

527. Rubber. This is the most important of the plastics, and the industry is now enormous, producing, per year, products valued at over a billion dollars. It is obtained from Malaysia, South America, and Africa. The South American Indians called it "cahuchu," whence our name caoutchouc (koochook), and the name *rubber* was given to it because it could erase or "rub" pencil marks.

The milky sap, or latex, is obtained by cutting through the

bark of the tree and catching the liquid in cups, as it drips from the tree. This latex contains caoutchouc, the rubber hydrocarbon, $(C_5H_8)_x$, apparently a condensation product of isoprene, C_5H_8 , in colloidal solution with a protein acting as protective colloid. It oxidizes slowly to a hard, brittle compound. It is soluble in most organic liquids like gasoline, benzene, carbon disulphide, and carbon tetrachloride. Rubber becomes useful only on coagulation. This used to be accomplished by the smoke of ordinary wood fires, which contains a little pyroligneous (acetic) acid, but, to-day, this primitive method is being displaced by the direct use of acids.

528. Vulcanization. Rubber was practically a useless article until Goodyear's discovery of vulcanization. Before vulcanization, rubber is sticky, inelastic, and soluble in organic liquids; after the treatment, it loses these properties and becomes strong and elastic, insoluble, and resistant to wear and weather. Vulcanization consists in adding powdered sulphur to the crude rubber and then after molding to the desired shape, heating to $150^\circ C$. without air. Exactly what takes place is not known, but it is certain that the rubber is still further coagulated. It may be an adsorption followed by chemical change, for the sulphur cannot be extracted by its solvent, carbon disulphide. Certain materials, called *accelerators*, are added to rubber to hasten the vulcanization. Zinc oxide is one of these.

529. How to make rubber goods. To make a rubber coat or shoe, the mixture of rubber and sulphur is spread on cloth, in the proper shape, and then vulcanized; that is, heated without air, to about $150^\circ C$. This is done by steam, under pressure. Soft rubber goods like sheets, tubing, and rubber bands contain less than 5 per cent sulphur, while hard rubber (ebonite) may have as much as 50 per cent sulphur.

White rubber contains chalk, clay, or zinc oxide, as filler; the red is antimony sulphide, the black is lampblack. Tires sometimes contain 50 per cent filler. Rubber erasers contain powdered pumice stone.



FIG. 294. Film made of celluloid.

530. Celluloid. Soluble guncotton (cellulose nitrate), pyroxylin, is mixed with camphor, and then pressed together at about 130°C ., giving celluloid, a solid solution of camphor in pyroxylin. On heating, it becomes soft, and can be molded. It comes on the market under many names, chief of which are imitation ivory, pyralin, and xylonite. It is used for photo films, combs, brush handles, auto windows, and innumerable other things, and it is an excellent article, but it has one drawback: it is very inflammable.

531. Cellulose acetate is used to make artificial silk and non-inflammable films. After molding to the desired shape, it is coagulated. In solution, it forms a lacquer for metals, and, applied to cloth, makes imitation leather.

532. Bakelite is made by warming together phenol, formaldehyde, and alkali. There are several varieties of bakelite: liquid and solid. The liquid may be applied like paint; the solid is plastic; but on heating either to 160°C ., under pressure, it sets to a transparent, colorless solid, almost like celluloid in appearance. (The yellow variety resembles amber.) It is non-inflammable, however, and is displacing celluloid for many purposes. It is a poor conductor of heat and electricity, infusible and insoluble. Pipestems, telephone receivers, billiard balls, phonograph records, umbrella handles, radio panels, and many other useful things are now made of bakelite.

533. Other plastics. Shellac is used to make cheap buttons. All sorts of gums are made into paints and varnishes, and these set after being

applied. *Linoleum* is made by applying a mixture of thickened linseed oil and granulated cork to a coarse cloth, like burlap. White lead, or any other powdered substance (red lead, litharge, and barium sulphate are often used) ground with linseed oil, becomes plastic and sets, due to oxidation of the oil. *Putty* consists of linseed oil and



FIG. 295. Automobile top, a use for cellulose acetate.



Bakelite Corp.
FIG. 296. Telephone made of bakelite.

whiting. Casein mixed with alkali, like borax or quicklime, makes a plastic mass, which sets, white and hard. This is called galalith. Some of these plastics, like glue, shellac, and casein, make excellent adhesives.

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 "Rubber," *Industrial and Engineering Chemistry*, vol. XVIII (1926), pp. 1104-1177.

QUESTIONS

1. Why are plastics useful?
2. How would you hasten the setting of cement?
3. Why does the dentist sometimes add salt to plaster of Paris?
4. How can most metals be made plastic?
5. Why is vulcanization of rubber necessary?
6. Name the steps (considering only the chemistry of the process) in making a rubber tire.
7. A fountain pen and an eraser may both be made of rubber. How do they differ?

CHAPTER XXXIII

EXPLOSIVES

534. History of explosives. The first mention of explosives in history is Greek Fire, used about 660 A.D. As far as can be gathered, it consisted of pitch, rosin, sulphur, and saltpeter. The first three are combustibles and the last, saltpeter, KNO_3 , is the



FIG. 297. Nobel, the inventor of dynamite.

source of oxygen, the supporter of combustion. This explosive was used to set fire to enemy houses and ships. The discovery of gunpowder was probably made in China centuries ago. In Europe, Roger Bacon, 1270, mentions the composition of black gunpowder in his writings, but it was in 1328 that Berthold Schwartz discovered the idea that stones or other projectiles could be propelled, with great force, by the use of this black powder, consisting of a mixture of niter or saltpeter (KNO_3), charcoal, and sulphur. In 1346, at the battle of Crécy, cannons were first used, and the doom of chivalry was at hand. Black gunpowder was practically the only explosive used during the middle ages. In 1845 Schönbein discovered guncotton, and since then improvements have followed rapidly. In 1847 Sobrero made

nitroglycerine, and in 1866 Nobel made it useful by converting it into dynamite.

535. Power of explosives. Explosives seem to be a source of enormous energy, but this notion is incorrect. One gram of combustible will produce the same energy whether it is burned slowly or rapidly. The difference is one of time, and the effect is due to inertia. If you receive a slow, forcible push, you will be moved, perhaps without pain. That same force applied rapidly becomes a blow, and you may not be moved, but the blow will be painful.

The energy developed by explosives has been measured. One gram of gunpowder liberates 600 calories, while one gram of coal in burning produces 8000 calories (explosives are costly sources of energy); but the gunpowder burns up in one second, whereas it takes the coal five minutes or more. At that rate, the gunpowder produces 180,000 calories *in the same time* that the coal produces 8000 calories. Some high explosives take only $\frac{1}{5000}$ second to liberate their energy, so that such a material would produce 900,000,000 calories while coal produced 8000 calories. This may explain the great destructive effect of high explosives. If a charge of low explosive is ignited on a solid surface in the open air, no shattering of the solid takes place, because the gases are produced slowly enough to escape against the resistance of the air. But a high explosive shatters a solid surface on which it is fired, because the products are freed so rapidly that even air offers sufficient resistance to their escape and allows the pressure on the solid to become great enough to shatter.

One cubic centimeter of gunpowder produces 500 cc. of products, chiefly gaseous, the solids producing the smoke. The gases include CO_2 , CO , and N_2 ; the solids, K_2CO_3 , K_2SO_4 , and K_2S . Now the 500 cc. of gases are heated to about 2500°C ., at which temperature they would occupy about 10 times their original volume, or 5000 cc. The space into which this volume is being jammed is 1 cc., the volume of the gunpowder before ignition; so that the pressure will be 5000 times the original atmospheric pressure of 15

lbs. per square inch ; that is, 75,000 lbs. per square inch. The same calculations applied to nitroglycerine gives a result of 1,500,000 lbs. per square inch, a pressure at which every known substance breaks. Furthermore the gunpowder develops its pressure during a whole second, while the nitroglycerine takes only $\frac{1}{5000}$ second.

536. Composition and properties of some explosives. Black gunpowder is a mixture of KNO_3 , called niter or saltpeter, with charcoal and sulphur. The proportions vary, but the following figures are a good approximation :

KNO_3	75%
C	15%
S	10%

The charcoal and sulphur are combustibles and the oxygen of the KNO_3 is the supporter of combustion.

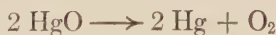
The ingredients are powdered separately and mixed while damp, using wooden implements. While still in this condition it is pressed into grains, small-grained powder being used for small arms, while for other purposes much larger granules are used. This compression brings the particles much closer to each other and shortens the period of combustion. Chemically it is very stable. Samples, centuries old, have been found unaltered in composition and properties. The tales of soldiers using gunpowder to salt the flesh of animals shot on the march are not unbelievable, since it is not poisonous, and since carbon and sulphur have no taste, the taste of the mixture being that of the niter, salty. It does not explode easily under a blow, requiring a spark or flame to ignite. It is, therefore, rather safe to handle. Its use has diminished very much with the discovery of other explosives, but it is still used as a primer, to ignite other explosives, in fuses, in blasting, where powerful shattering is not desired, as in quarrying and mining, and in fireworks and shells for small firearms.

537. Firing an explosive. Gunpowder and other low explosives are fired by a spark produced by a cap, fuse, or train of powder. In

the open air, it burns rapidly, with a huge flame, but very little sound, and no rending or destructive effect. The slow match for firing cannon consisted of loose rope soaked in a solution of lime-water and niter, so that it would carry a spark, but not too rapidly, about five inches an hour. The quick match was made of cotton wick soaked in a solution of gunpowder and gum, finally strewn with dry gunpowder. When dried, this fuse burned about three inches per second. High explosives are fired by detonators, which in turn are fired by fuses, or the detonator may be fired electrically. Electric firing has the advantage of setting off all the charges at once. Fuse firing may be regulated so that the charges are fired separately and counted, insuring that all have been fired. This is safer, of course.

538. Detonation. High explosives are fired by detonation; low explosives by ignition. Detonation is brought about by the explosion of a small amount of some other explosive, called a detonator. This seems to cause a simultaneous reaction in all parts, supposed, by Abel, to be due to vibrations produced in the explosive, similar to those of the detonator — like sympathetic vibrations. If a crystal of potassium chlorate is broken in contact with phosphorus, a violent explosion occurs; if in contact with sulphur it is less violent, and if sugar replaces sulphur, it merely burns. Perhaps the breaking of the crystal of KClO_3 sets free a very little oxygen, which, combining with phosphorus, produces heat sufficient to continue the reaction. A blow will do the same, and detonation produces great pressure, similar to a blow. All good detonators contain a heavy atom as in the fulminates of mercury and silver, increasing the similarity to a blow.

539. Decomposition of unstable compounds. When elements do not readily combine, their compounds are, of course, easy to decompose. Mercury and oxygen combine, to form the red powder, HgO , but with difficulty. For that reason it is easy to decompose HgO :

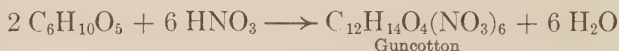


Nitrogen is even more inert than mercury. It follows, then, that nitrogen compounds should be rather loosely joined, unstable, and, therefore, decomposition will be easy. Nitrogen chloride, for instance, when touched, will explode.



The gases produced are very much increased in volume by the heat developed during the decomposition, which is very rapid, and we have the conditions for explosion. Here it is a decomposition, which, started, will continue and even get faster. Most of our explosives are of this nature; they are unstable compounds; in fact, nitrogen compounds.

540. Nitrocellulose or guncotton. When pure cotton fiber is treated with a mixture of nitric and sulphuric acids, it looks just the same as at first. But if a flame is applied, it burns up instantaneously, with no smoke or ash and no sound. A blow or friction will cause a violent explosion. The guncotton is compressed after being pulped, and sawed and bored, while wet, into shapes required. Whereas cotton is practically insoluble in everything, the nitrocelluloses dissolve in acetone to a viscous liquid. The lower nitrates dissolve also in alcohol and ether mixture, forming "collodion," while the higher nitrates are insoluble and it is to them that the name "guncotton" is usually given. Although insoluble, guncotton becomes plastic in ether and alcohol, and in this condition it is rolled into sheets and cut into small pieces. This is *smokeless powder*. The lack of smoke is due to the fact that all the products of combustion are gases, for only solids can produce smoke. The equation for the formation of a nitrocellulose can only be written if we assume the formula of cellulose, which is not known exactly. We then have for the formation of the hexanitrate

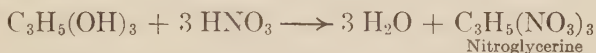


The sulphuric acid does not appear in the equation, being used to remove water (dehydrating agent), one of the products of the

reaction; in this way constantly upsetting equilibrium and increasing the speed of the reaction. The hexanitrate contains more than enough oxygen for complete combustion.

The reason why a compound like nitrocellulose will explode so much more rapidly than a mixture like gunpowder is, probably, because in a compound the atoms which are to combine with each other are in much closer contact than they can possibly be in a mixture. Guncotton in a moist condition is very safe to handle and still it will explode with a priming charge of dry guncotton. Guncotton is used to make torpedoes and mines and smokeless powders, but the large volume of carbon monoxide produced makes it unfit for industrial use except in open air.

541. Nitroglycerine. This is prepared from glycerine in a way very similar to the preparation of nitrocellulose, according to the equation:

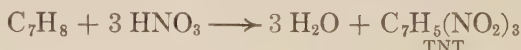


Sulphuric acid is again used to dehydrate and so increase the speed of reaction. Nitroglycerine is a yellowish, odorless, oily liquid, insoluble in water, but soluble in organic solvents like acetone, ether, etc. It freezes rather easily, at 12° C., and the solid is even more difficult to handle than the liquid. It must therefore be thawed, before use, by careful heating in steam rooms. Mixing with dinitroglycerine prevents freezing. In the open air, it may be made to burn, harmlessly, but confined, it explodes at 180° C., the speed being very high, 7500 meters per second. The chief products of the explosion are CO₂, H₂O, and N₂. It will be noticed that nitroglycerine contains more oxygen, within itself, than is necessary to burn up the combustibles completely. Advantage is taken of this fact to produce an even more powerful explosive, by mixing with it about 10 per cent of nitrocellulose. On cooling, this sets to a transparent jelly called *blasting gelatin*. Its rate of detonation is 7500 meters per second, and the temperature about 5000° C. The forty per cent mixture is quite safe to handle.

Nitroglycerine, alone, is not much used, except in smashing oil wells. Whenever, on boring a hole to the oil stratum, the oil flows very slowly, the cause may be that it is being impeded by rock. In such a case, a charge of nitroglycerine is lowered and fired, so as to smash up the rock and allow the oil to run together. But for most purposes a liquid is difficult to handle, because it cannot be kept in place. At one time nitroglycerine was used in mining on the Rand, in South Africa. Holes were drilled by hand, and when ready, a little nitroglycerine would be poured into each hole, and at the end of the shift, all holes were fired at once. A number of serious accidents, however, put a stop to this practice. It seems the liquid explosive would run along cracks in the rock and when, in drilling new holes, this was struck, a violent explosion would follow. When the liquid is absorbed in fuller's earth, we get dynamite, a plastic solid, easily handled and quite safe. Dynamite is fired by detonating cap, set off electrically or by a fuse. It is used in mining and blasting.

542. Picric acid. Trinitrophenol is made from phenol (carbolic acid, C_6H_5OH) by the action, first, of sulphuric acid and then of nitric acid. It is a yellow solid, once much used in dyeing, whence the nickname of "canaries" for some of the workers in dye houses. Picric acid itself is quite safe to handle, but some of the picrates are very sensitive to blows, notably lead picrate. The acid may be easily fused and in this condition poured into shells, which is its chief use. In this form it is known as lyddite.

543. TNT or Trinitrotoluol is one of the recent high explosives first introduced in the late war. It was discovered more than 50 years ago, but it is so difficult to detonate (pieces of it can be ground up or broken up with a hammer without explosion) that it was known only in the laboratory. It is made by direct action of nitric acid on toluol, C_7H_8 , a coal-tar distillate resembling benzene, according to the equation



It looks something like picric acid, yellow crystals, is insoluble in water but soluble in organic solvents. It melts at about $80^{\circ}\text{C}.$, which makes it easy to fill shells. It is poisonous to handle, and many workers were badly affected by it, during the war.

On explosion it yields 70 per cent CO , a little CO_2 , and for the rest N_2 , H_2 , and C . To overcome this, oxidizing agents are frequently added. Its rate of detonation, about 7000 meters per second, its chemical stability, and its great ease of handling, make it almost an ideal explosive for military purposes. It is detonated by the explosion of a small quantity of some other explosive. In time shells, the fuse, whose length can be varied, is lighted by a fulminate primer, set off on firing the shell from the gun. The explosion of the shell then follows, a definite time after its discharge from the gun. A percussion fuse, on the other hand, will start the explosion only when the shell hits some object. A shrapnel shell contains only a low explosive and is filled with bullets. Bombs and hand grenades contain high explosive, which need not be quite so insensitive as for shells, since they are not fired from a gun.

544. Permissible explosives. In gassy or dusty coal mines, the danger of explosion may be ever present. Artificial ventilation may reduce the proportion of combustible gas below that necessary for an explosive mixture, and the use of safety lamps may prevent the mixture reaching its ignition point, if it is explosive. But what about blasting? When black gunpowder was the only explosive used, mine explosions were very frequent; but a study of the problem has revealed some interesting facts, and reduced the danger considerably. It was found that in igniting an explosive mixture, the temperature is not the only factor, but the duration of that temperature. As already stated, natural gas and air explode at 650° with a flame of one tenth second duration. Black gunpowder gives a flame of $2200^{\circ}\text{C}.$, for one second, while 40 per cent dynamite gives a flame of 2900° , but for only $\frac{1}{2500}$ second. Permissible explosives, as they are called, are those safe for use in gassy coal mines. They are usually made of dynamite doped with water,

either as liquid or as water of hydration in alum or magnesium sulphate. Wood pulp, corn meal, and flour are also used as dopes. These have the effect of reducing the flame temperature, and since the time of explosion is very short, such explosives are quite safe.

QUESTIONS

1. Explain why a high explosive shatters.
2. What is smokeless powder? Why smokeless?
3. Name some peace-time uses of explosives.
4. Name some engineering enterprises which could not have been carried out without explosives.
5. How did Germany get its explosives during the war? (It was cut off from Chile, the source of nitrates.)
6. Why are nitric and sulphuric acids used in making modern high explosives?
7. Suppose you had to blast out a tree stump. What explosive would you use? Why?
8. In the explosion of one kilogram of nitroglycerine, what is the total volume of gases produced?

OPTIONAL QUESTIONS

9. Consult any encyclopedia, and write, in your own words, an account of the history of gunpowder.
10. Why are there so few natural compounds of nitrogen? Use your answer to explain why water or hydrochloric acid is a stable compound.

CHAPTER XXXIV

PHOTOGRAPHY

545. It has long been known that many remarkable changes are brought about by the action of light. From time immemorial, laundresses have spread their linen in the sun to be bleached, and they recognized this effect of light more definitely by protecting their colored fabrics from the direct sunlight. The most important chemical action of light is the changing of carbon dioxide and water (in the presence of chlorophyll) into starch (photosynthesis). It has been shown that chemical action under light increases as we approach the violet end of the spectrum, and that, in fact, ultraviolet light has the greatest effect. (Ultraviolet light is composed of blue rays that human eyes cannot detect.) It is this light that causes sunburn, and living things deprived of it do not thrive. Window glass is not so permeable to ultraviolet light as quartz, and animals and plants must be exposed to direct sunlight or suffer in health. Exhaustive experiments with chickens kept in coops with glass windows showed that they were poor in health compared with chickens grown in coops with quartz windows; whereas these latter were equal to another brood kept in the open. We now have special types of glass which permits ultraviolet light to pass through. Ultraviolet light has been used, with success, in curing rickets and tuberculosis of the joints.

Light of different wave lengths (colors) produces chemical effects in varying times, as follows: violet, 15 seconds; blue, 30 seconds; green, 40; yellow, 300; red, 600. Light producing chemical change is called actinic. It has long been known that iron and silver compounds are easily affected by light, as for instance the black stains produced by silver nitrate (lunar caustic).

546. Blue printing. When ferric compounds mixed with reducing agents are exposed to sunlight or other ultraviolet light, ferrous compounds are formed ; as, for instance,



If this is now treated with potassium ferricyanide, the blue precipitate, ferrous ferricyanide, called Turnbull's blue, is thrown down. In actual practice, blue-print copies of mechanical drawings are made according to the following steps :



FIG. 298. Floor plan made on blue-print paper.

1. Blue-print paper made by coating a hard finish paper with a solution of ferric ammonium citrate ; that is, a ferric compound containing reducing agents.

2. Drawing made on tracing linen with India ink.

3. Exposure of coated paper under drawing in sunlight, arc, or mercury vapor light.

4. Reduction of ferric salt to ferrous wherever light penetrates.

5. Washing the blue print.

Water brings about the reaction between ferrous salt and potassium ferricyanide, and the insoluble blue appears. Wherever the light did not reach, the original salts, soluble in water, are washed off, so that we get white lines on blue paper.

One may write on the blue part of the print with sodium hydroxide, which changes the complex ferrous ferricyanide to $\text{Fe}(\text{OH})_3$, from which it ought to be evident that blue prints cannot be washed in alkaline solutions, like soap or soda.

547. Silver salts in photography. In the sixteenth century, the alchemists noticed the blackening of horn silver, AgCl , on exposure to light; but it was not until 1800 that Wedgwood succeeded in producing a picture by this means. This was not, however, fixed, but resembled our photographic "proofs"; so that on further exposure they darkened, and soon disappeared. In 1839, Daguerre perfected his process of making permanent photographs: a metal plate, coated with silver iodide, was exposed in a camera, and this plate was then treated with vapor of mercury, which attached itself to the particles of metallic silver in the exposed picture. These were called daguerreotypes. Each exposure produced only one photograph.

The modern process of photography begins like that of Daguerre, consisting of these steps:

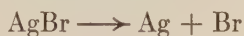
1. Making the film.
2. The exposure.
3. Development.
4. Fixing. This produces the finished negative.
5. Printing and toning, to produce positives.

These steps will now be taken up and discussed in order.

548. Making the film. An emulsion of silver bromide is produced by adding, slowly, silver nitrate solution to a solution of sodium bromide mixed with warm gelatin. The gelatin protects the precipitate of silver bromide, AgBr (colloidal), so that its particles are exceedingly small, and it is "ripened" by keeping the gelatin warm, and allowing the particles to grow to the proper degree of dispersion; for it is found that silver bromide is more sensitive to light when its particles are coarse but still in the colloidal condition. A little silver iodide is usually mixed with the silver bromide. This is all done in a "dark room," where the illumination is produced by red light.

When the emulsion is ripe, it is spread on glass plates or celluloid sheets, called films, and allowed to set, the gelatin acting as an adhesive. The film is now ready for exposure.

549. The exposure. We shall not, here, enter into the mechanism of the camera, except to state that the film is held in such a position that it can be exposed to light, reflected from the object, for the proper length of time. The light which is most effective contains ultraviolet rays present in sunlight, arc, mercury vapor or flashlight (mixture of $\text{Mg} + \text{KNO}_3$). Light of other colors has less effect on the silver bromide. Just what chemical action takes place is not well understood, but it seems to be agreed that a reduction of AgBr to Ag takes place, wherever light strikes it. This may be represented by the equation



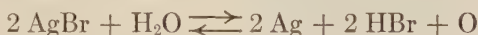
The gelatin combines with the Br , and so, removing it from the system, hastens chemical action. Silver iodide is more unstable than silver bromide, which accounts for its greater speed, but bromine is more active in combining with the gelatin than iodine (these two factors balance); but if to the emulsion is added some substance which combines with iodine, such as an alkali, the silver iodide is most sensitive. The exposure has produced, then, a "latent" image of colloidal silver; that is, nothing can be seen, but it is capable of development (latent means hidden).

550. Development. The exposed film is now treated, in a dark room, with a "developer." This consists of a reducing agent, hydroquinone or pyrogallol, mixed with an alkali. And why are these used? The reducer continues the production of more colloidal silver, started during the exposure. And why is not the unexposed silver bromide reduced? It is, but not so quickly. We shall understand this apparent contradiction by comparing this reaction with combustion. If a combustible is heated slowly in air, its rate of oxidation is at first very slow; it accelerates as the temperature increases, until a point is reached (kindling temperature) where oxidation suddenly becomes enormous. Now, then, if we heat one combustible, and later start heating another, and thereafter continue heating both, no apparent change takes place

until one of them suddenly bursts into flame, and if we now stop heating, the other will not take fire. Likewise, in development, we treat with a reducer two kinds of silver bromide, AgBr: one whose reduction has been started, and the other untouched. Naturally, the light-struck AgBr reaches its critical point sooner, and one can notice the image almost leap out. If the development is continued, the rest of the AgBr reduces, and the entire film darkens. This is called overdevelopment.

The developed film now contains gelatin in which is suspended finely divided black metallic silver; heavier, where the light was strong. Wherever the light did not affect it, of course, the silver bromide remains unaffected.

To understand the function of the alkali in development, consider the reversible equation



We have learned, in the chapter on Chemical Action in Solution, that, because of mass action, this reaction can be accelerated to the right, by removing one or more of the products. Now, the oxygen is taken care of by the reducing agent, while the HBr is neutralized by the alkali, usually Na_2CO_3 , producing NaBr. Na_2CO_3 is much better than NaOH, because too much alkali (that is high C_{OH^-}) causes swelling and softening of the gel; whereas Na_2CO_3 gives hydroxyl ion, but not in high concentration; it acts as a reservoir of hydroxyl ion.

Sodium sulphite, Na_2SO_3 , is added to developers to preserve them: it takes up oxygen from the air, which would otherwise oxidize the reducing agent in the developer and weaken it. It also hardens the film.

It is told that development was discovered by Talbot in a strange manner: his cat upsetting a jar of nutgalls extract (a reducing agent) on some exposed paper covered with silver chloride.

551. Fixing. If the developed film were to be taken out into the light, it would turn entirely black, due to the presence of unre-

duced silver bromide. This silver bromide must first be removed, and that is called fixing. For this purpose, the film is first washed in water, to remove developer, and then immersed in a solution of sodium thiosulphate (or hyposulphite; "hypo" for short),

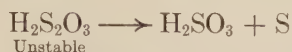
$\text{Na}_2\text{S}_2\text{O}_3$. This dissolves AgBr and other salts of silver, and leaves on the film only the gelatin with dispersed silver, a finely divided black deposit. This can no longer be affected by light, and so we have a permanent, "fixed" image, called a negative. It is so called because colors are reversed: a light colored object, reflecting much violet light, affects the silver bromide very much, and so produces a heavy deposit of silver, which is black. A dark object does not reflect much light, and so the silver bromide is unaffected, and ultimately washed off by hypo, leaving transparent gelatin and film. This is shown in the accompanying photographs of negative and positive.



FIG. 299. Photographic "negative" and "positive"

If developer is carried over into the hypo, it will ultimately become oxidized by air, and brown stains result. This is prevented by the presence of Na_2SO_3 , which itself takes up the oxygen. The sodium sulphite serves another purpose. To harden the gelatin (softened by alkali in development), acids are used in the fixing solution. Now, a highly ionized acid acts on hypo to pro-

duce thiosulphuric acid, $\text{H}_2\text{S}_2\text{O}_3$, which is unstable, breaking up according to the equation



and according to the Law of Mass Action, the addition of Na_2SO_3 would force this reaction to the left, and so prevent it (common ion effect). Another way to prevent this, is to use only acids of low concentration of hydrogen ion, such as citric, tartaric, and acetic acids.

552. Washing. Fixing solution must not be used indefinitely, because when fairly saturated with silver, if some of it is carried out (due to insufficient washing) by film or paper, Ag_2S stains will ultimately be formed by hydrogen sulphide in the air.

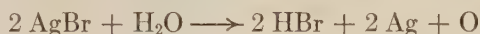
Washing should not be done with warm water, because the gelatin swells and this hinders diffusion. A simple test to determine when washing is complete is to use a little potassium permanganate in the final wash water. This loses its pink color in hypo, due to reduction.

553. Printing. The negative may now be used to produce any number of positives, called prints. This process is very much like the original exposure: a piece of paper coated with sensitized emulsion is exposed to light under the negative. Light penetrates the negative where no silver has been deposited, and produces a latent image on the paper. This is developed and fixed, producing the permanent picture.

554. Toning. The colors of such prints, due to colloidal silver, are not so pleasing as those of gold and platinum — nor so permanent. Replacement of gold or platinum for silver in the print may be brought about by immersing it in solutions of gold or platinum compounds. These produce black, green, or blue tones. Rate of replacement is important, since slow action produces red colloidal gold, whereas rapid replacement gives blue gold. Purples or browns are also possible. Sepia prints may be produced by converting the silver image to silver sulphide, which gives brown

to black deposits. The use of dyes in toning is an interesting development: the silver image is changed to silver iodide, and this silver iodide acts as a mordant for basic dyes. In this way, prints of any desired color may be made.

555. Some other points explained. *Restrainer.* Potassium bromide, KBr, is frequently used to slow down the process of development. A consideration of the equation will explain:



Addition of KBr increases the concentration of bromide ion very greatly, decreasing the ionization of AgBr and HBr; and since speed of reaction depends upon degree of ionization, it is apparent that the reaction must slow down.

Accelerator. Vice versa: if we use something which can remove one of the products of the reaction, the speed will be increased. Sodium carbonate, Na_2CO_3 , can neutralize HBr, and so accelerate the change.

Reduction. This term is unfortunate, because it is used in one sense in chemistry, and in a totally different sense in photography. In chemistry, reduction usually implies gain of electrons; whereas, in photography, it means weakening, in cases of overdevelopment. Now, in such a case, there is too much silver on the negative, so that reduction means removal of silver. If the negative is treated with a solution of potassium permanganate and H_2SO_4 , silver is oxidized to silver sulphate, which is soluble.

Intensification. This means just the reverse of reduction; that is, increasing the opacity by depositing additional silver (or mercury) on the negative. For instance, immersing the negative in a solution of mercuric chloride, we get



The negative is then redeveloped, when AgCl and HgCl both reduce to metallic Ag and Hg, thus depositing twice the original amount of metal.

556. X-ray photography. The discharge from X-ray tubes consists of very short light waves which affect photographic plates, while it differs from ordinary light in the fact that it penetrates organic material and even light metals; only heavy metals, like lead, being opaque. You may recall in the chapter on "What Is Matter," that the nucleus is rather concentrated, or dense, and this is especially true of the heavy metal atoms. When the X-rays are directed at organic material, the atoms of C, O, H, and N are not dense enough to interfere; whereas the metallic atoms may be hit, causing a blur, as when one looks through ground glass.

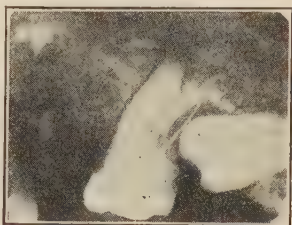
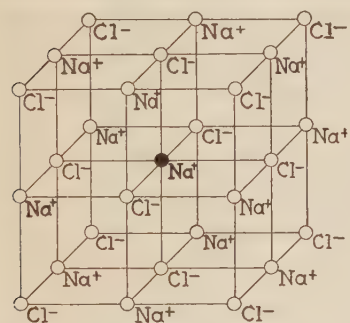


FIG. 300. X-ray photograph.

To take an X-ray photograph, the object to be photographed is placed between the X-ray tube and the film, folded in black opaque paper. The rays pass through all objects but dense ones — like bones — and so a shadow photograph is produced on the film.

X-ray photography has advanced so rapidly that at the present time more films are used for this than for portraits. One is more

apt to have a photograph taken of his teeth than of his face.



rays which penetrate a thin section of the crystal will show the position of the atoms in the molecules that make up the crystal.

The accompanying diagram gives some idea of the structure of the sodium chloride crystal.

557. Analysis of crystals. Recently, Laue, the Braggs, and others have made use of X-rays to study the arrangement of atoms in a molecule, and of molecules in a crystal. The same atom always has the same effect on the rays, and a photograph made by the

SUPPLEMENTARY READING

- Sir Wm. Bragg, *Concerning the Nature of Things* (Harper & Bros., 1925).
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H. E. Howe and F. M. Turner, *Chemistry in the Home* (Charles Scribner's Sons, 1927), chap. xii.
H. E. Howe, *Chemistry in Industry* (Chemical Foundation, 1924), vol. I, chap. xviii.
J. Southworth and T. L. J. Bentley, *Photographic Chemicals and Chemistry* (Sir Isaac Pitman & Sons, 1927).

QUESTIONS

1. Why is a red light used during the development of films?
2. Explain the presence of white lines on a blue print.
3. Write equations for the reduction of a ferric salt by oxalic acid, in blue printing.
4. Write equations for the development of a photograph.
5. Why is fixing necessary?
6. What is ripening?
7. What two purposes are served by the gelatin on a film?
8. Of what use is acid in the hypo?
9. Define positive; negative; toning; "hypo"; intensifier.
10. Why does a soluble bromide "restrain" the development?
11. Make a diagram showing proper positions of X-ray tube, film, and subject, for taking an X-ray photograph.
12. Explain why it is not just as good for the health to sit on a glass-enclosed porch as in the open, during the day.
13. Why cannot salts of metals other than silver be used in photography?
14. What is the chemical action during exposure?
15. Which of the salts of silver is the most sensitive to light, and why?
16. How might the study of history have been affected, if photography had been invented, in ancient times?
17. Show in what respect photography has influenced the development of astronomy, medicine, the movies, the newspaper.
18. Write an equation to show transfer of electrons during the reduction of a ferric salt in blue printing.
19. Write the electron equation for "development."

CHAPTER XXXV

PAINTS AND VARNISHES

558. Surface protection. In the preservation of our material possessions we must foresee the possible changes they are liable to undergo. Now, all material may be subject to chemical change by new combinations, and it may also change its condition gradually from colloid to crystal. What material will therefore be best to build a house? For lasting quality, for resistance to the action of natural weathering agents, surface rock will be best; because having successfully withstood the action of nature for many centuries, it ought to last a few more decades. Furthermore, most rock is crystalline, so no more changes in that direction need be feared. Quarried stone, however, may have been taken from far below the surface, and may be attacked by air and water. Brick is rather inert, for the same reason, — it is made of clay and sand, both of which exist in nature, unchanged, for long periods of time. Metals are crystalline already, and therefore do not undergo a change in condition; but most metals (except platinum, gold, and other so-called “noble” metals, which occur free in nature) are affected chemically by air and water. Wood, textiles, and other organic matter are, of course, colloidal, and therefore have changes in condition as well as in chemical composition to undergo.

Now, in considering possible chemical changes, we must remember that two substances react only where they touch, at their surfaces of contact; so that, for instance, a piece of zinc, inclosed in rubber, would not set free hydrogen, if immersed in sulphuric acid. It has also been noticed that chemical action is more intense at points and edges; that is, where the specific surface is great. Notice how bubbles arise always from scratches or points in the vessel. Put a piece of metal in acid and notice how most of the

bubbles arise from the edges. Then scratch the surface, and now bubbles arise from the scratch. These facts explain why smooth material is not so readily attacked as rough; and should concentrate our attention on protecting the surface as a means of protecting the entire article. Some surfaces need no protection, like rock, as already mentioned; some are sufficiently protected by smoothing, but most things will need surface protection, and that means not only those surfaces that can be seen, but all parts that come into contact with air, water, or other chemical reagents.

From time immemorial, wood has been protected from insects and bacteria by charring. So effective has this proved, in some cases, that some of the timber which was used by Caesar to build a bridge across the Rhine is still in well-preserved condition. Creosote and zinc chloride are among the substances used, nowadays, to protect wood.

Metallic protection is given to many articles, particularly of iron, but these have been referred to under their appropriate heads. Non-metallic surface protective agents are called paints, varnishes, and enamels.

559. Paint. Most paints are made of material so coarse that rapid settling takes place, and only lately has any attempt been made to apply the principles of the colloidal condition — as in the use of a protective colloid to aid suspension; rather have they used viscous (thick) mediums, which provide only a lame remedy. A paint, then, consists of two essential parts: a liquid, called the vehicle, or medium, and a solid, called the base. Besides these, one or more of the following are used:

SUBSTANCE	REASONS FOR USE
Thinner	secures proper consistency
Drier	hastens oxidation and coagulation
Pigment	provides color
Extender	dilutes pigment; etc.

560. Drying oils. A paint differs from a dye in that dyes are true or colloidal solutions, and they are rendered fast by adsorption on fibrous material. Very little protection is afforded by dyes, whereas paints form a comparatively thick layer of impervious gum. The most important ingredient of the paint from this standpoint is, therefore, the vehicle. No sharp distinction can be drawn between the thinner and the vehicle, in some cases, as in whitewash, but if we accept the idea that the vehicle must not evaporate completely, but must coagulate to form a protective gel, then there are only very few such liquids: the so-called drying oils. If we expose, on plates, small quantities of various oils, we find the results shown in the table:

OIL	RESULT OF EVAPORATION TEST
Olive	remains liquid — unchanged
Gasoline	evaporates completely
Kerosene	evaporates partially — residue still liquid
Turpentine	evaporates, but leaves gummy residue
Linseed oil	does not evaporate; gels
Corn oil	gels partially

Linseed and corn oil, and, to some extent, turpentine, are called drying oils. The most important drying oils are linseed, poppy, and Chinese wood oils, and of these, linseed is the most common, in this country. These are sometimes mixed with fish oil, rape, cotton, hemp, corn, rosin, and petroleum oils, which may be regarded as adulterants.

Linseed oil is obtained from completely ripened flaxseed, either by pressing or by dissolving with gasoline. In the latter case, the gasoline is evaporated, recovered, and used again. Argentina, United States, Russia, and Canada are the chief producers of flaxseed. Unfortunately, it is not possible to obtain both flax (to make linen) and linseed oil from the same plant; because the fiber is taken while the seed is not yet ripe. The oil is refined with sul-

phuric acid, which coagulates certain constituents and prevents thickening in spots. Boiled oil is raw oil incorporated by heating, not boiling, with a small amount of lead or manganese linoleate.

Linoleum or floor oilcloth consists of a sheet of canvas covered with hardened linseed oil and cork pressed on by rollers. Putty is a mixture of linseed oil and whiting. Rosin incorporated with Chinese wood oil makes a vehicle considered by Toch as superior to linseed oil.

561. Bases. The material next in importance is the base or solid constituent of the paint. The most important of the bases is white lead, but there are many others, some white and some colored: all materials which can be ground with oil to make a complete paint. Some of these with their colors are shown in the table.

NAME	COLOR	COMPOSITION
White lead	white	$\text{Pb(OH)}_2 \cdot 2 \text{PbCO}_3$
Zinc white	white	ZnO
Barytes	white	BaSO_4
Lithopone	white	$\text{ZnS} + \text{BaSO}_4$
Lampblack	black	C
Graphite	black	C
Yellow ocher	yellow	$2 \text{Fe}_2\text{O}_3 \cdot 3 \text{H}_2\text{O}$
Indian red	scarlet	Fe_2O_3
Venetian red	brick red	$\text{Fe}_2\text{O}_3 + \text{CaSO}_4$
Red lead	scarlet	Pb_3O_4
Burnt sienna	mahogany	$2 \text{Fe}_2\text{O}_3 \cdot 3 \text{H}_2\text{O} + \text{MnO}_2$
Umber	yellow brown	$2 \text{Fe}_2\text{O}_3 \cdot 3 \text{H}_2\text{O}$ and more MnO_2
Chrome green	emerald	$\text{PbCrO}_4 + \text{Prussian blue}$
Chrome yellow	yellow	PbCrO_4

562. Manufacture of white lead. White lead is a basic carbonate, made essentially by the treatment of lead acetate with carbon dioxide. In the French process, this is done by passing a

stream of carbon dioxide into the solution of basic lead acetate. In the old Dutch process, by which the great quantity is made, perforated pieces of lead are piled in pots with vinegar and tanbark. The vinegar produces lead acetate, which is acted upon by the carbon dioxide, produced by the decay of the tanbark. The process takes three months, but is supposed to produce a better product, with smaller particles.

563. Zinc white. Zinc white, ZnO , is made by burning metallic zinc. Its particles are smaller and whiter than those of white lead, and so, forming a very thin layer, more coats must be used, if it is used alone. For this reason it is seldom so used, except in white enamel paint, but is mixed with coarser-grained barium sulphate or white lead. It produces a white, hard, smooth surface, which does not blacken, but which cracks and peels. When fresh, it keeps off dust very well, but it is hard to repaint.

	ADVANTAGES	DISADVANTAGES
WHITE LEAD	cheaper than zinc white easy to repaint great covering power	blackens on exposure (forms PbS , black) soft finish (holds dirt)
ZINC WHITE	whiter than white lead hard finish does not blacken	peels hard to repaint more coats needed to cover

Apparently each of these has its advantages, and a combination may be so effected as to retain most of the advantages of each. In general, for outside work, it is preferable to use two thirds zinc and one third lead, while for indoor work, two thirds lead and one third zinc is better. Also the first coat should be all lead, the coarser particles being better as a foundation. Both white lead and zinc white are ground in oil to the consistency of putty and marketed in cans or pails, and must be mixed with more oil and turpentine

before being used. The white lead is slightly yellow after being mixed with the yellow linseed oil and requires a little blue pigment to whiten it. This should be Prussian blue, not ultramarine, because the latter contains sulphur which blackens the lead in a short time.

564. Pigments. If the color of the base is satisfactory, the paint is complete; but otherwise a pigment of the desired color must be added. The pigments used in oil paints are all insoluble materials and should be in the colloidal condition. But up to the present, the only method of increasing the degree of dispersion has been mechanical grinding. A pigment should be inert, chemically, and, in fact, many of them are made of natural minerals. The fact that they have maintained their color while subjected to "wind and weather" for an incalculable period of time recommends them for use in outside paints; and, except for good reason, such a material should always be chosen. Some of these materials with their colors are here given:

COLOR	PIGMENT	MINERAL FROM WHICH MADE
black	graphite	graphite
yellow	yellow ocher	limonite
yellow	chrome yellow	crocoisite
yellow	king's yellow	orpiment (arsenic trisulphide)
orange-red	realgar	realgar (arsenic monosulphide)
red	Venetian red	hematite + gypsum $\text{Fe}_2\text{O}_3 + \text{CaSO}_4$
red	Indian red	hematite
vermilion	red lead	minium (red oxide of lead)
brown	umber	limonite + pyrolusite
green	chrome ocher	Cr_2O_3
blue	ultramarine	lapis lazuli

We do not mean to imply that chemically prepared colors are necessarily inferior, but certainly some of them do not stand up

against weather, and one must know before choosing. Some of the prepared colors that are good are :

COLOR	NAME	COMPOSITION
black	lampblack	C
black	asphalt	C + hydrocarbons
white	white lead	$\text{Pb}(\text{OH})_2 \cdot 2 \text{PbCO}_3$
white	zinc white	ZnO
blue	Prussian blue	ferric ferrocyanide
green	verdigris	$2 \text{Cu}(\text{C}_2\text{H}_3\text{O}_2)_2 + \text{CuO}$
green	Paris green	$\text{Cu}(\text{C}_2\text{H}_3\text{O}_2)_2 + 3 \text{Cu}(\text{AsO}_3)_2$
green	Scheele's green	HCuAsO_3
orange	chrome orange	basic lead chromate
red	vermilion	HgS
red	chrome red	basic lead chromate

For china painting, the colors must not be affected by heat, and for this purpose they use the same materials as those which give colored borax beads.

COLOR	OXIDE USED
blue	cobalt
blue-green	copper
green	chromium
yellow to red	iron

565. Thinners. The common thinners used for paints are oil of turpentine, petroleum oils (similar to road oils), coal tar, naphtha, and denatured alcohol. Amyl alcohol and amyl acetate are used for collodion paints. For oil paint the best, by far, is oil of turpentine, $\text{C}_{10}\text{H}_{16}$. It has a boiling point of 150°C ., and is obtained from the destructive distillation of pine sap (the residue being resin). It dissolves wood resins and therefore is used in the first coat, to penetrate the wood. It causes paint to dry "flat" (dull). Petroleum distillates are poor substitutes for

turpentine. When used to thin paint, turpentine dissolves the linseed oil, but its evaporation is slow and this seems to allow time for the paint to gel, whereas with more rapid evaporation, as in the case of petroleum oils, the paint is more powdery.

566. Driers. The setting process in paint is accompanied by an oxidation of the oil, which starts slowly, like most reactions, and accelerates after a while. By boiling up some linseed oil with certain materials, chiefly oxidizing agents, like lead and manganese oxides, we get driers, which when added in small quantity to paints cause them to dry much faster (catalysis). Red lead ground in oil sets rapidly and is used by plumbers to make tight connections in pipes. Boiled linseed oil is not really "boiled," but is heated with a little drier, and this causes more rapid drying of the paint. At the same time driers cause paints to set hard and to crack; their use therefore should be limited.

Japan drier is a solution in turpentine of lead or manganese linoleate, and is the drier commonly used.

567. Extenders. These are used to dilute or adulterate pigments, but sometimes their use is warranted for a special reason. The chief extenders with uses and properties are covered in the table.

NAME	PROPERTIES	USE
Barytes (BaSO_4)	not much covering power; colloidal material better than natural crystals	mix with ZnO for body; adulterant of white lead
Silex (SiO_2)	rough; sharp	filler. Gives "tooth" to paint
Asbestine (magnesium silicate)	prevents settling	fireproof paints
Kaolin	prevents settling	filler
Gypsum	modifies colors	with Venetian red
Chalk (whiting)	basic	neutralizes acidity; calcimine; putty

568. Varnish. A varnish is usually a true solution of a resin or gum in a thinner. A drying oil is nearly always added, so that a varnish may be looked upon as a paint without a base but with the addition of a gum. The same thinners and driers are used as in paints. The outside varnishes, called spar varnishes, contain more oil, and it is usually Chinese wood oil.

The Resins. Copal, a fossilized gum, obtained from the exudations of tropical trees, is the best of the gums; it makes a hard, glossy finish, but is not too brittle. Amber and lac rosin also belong to this class. Juniper, sandarac, mastic, and dammar produce a rather soft finish but not elastic, while benzoin and elemi produce soft, elastic gums. Collodion is also used to make varnishes. Rosin, left after distillation of turpentine, is the chief ingredient of cheap varnish. It turns white in water, produces a brittle film, but when mixed with other gums and Chinese wood oil, particularly for inside work, may be quite good. Shellac comes from India. There are two varieties: orange shellac, the natural, and bleached shellac. It is dissolved in alcohol, which, evaporating rapidly, makes a quickly drying varnish; but it is brittle and does not resist the action of water. It is excellent for under coats, however, producing a brilliant finish. Mastic and sandarac are frequently used for the same purpose. Flatting varnishes have a wax mixed with the gum. Chinese wood oil varnishes frequently dry flat. *Lacquers* are very thin spirit varnishes. Black varnishes are usually pitch or asphalt in turpentine or benzene.

569. Enamels, more properly called enamel paints, are mixtures of oil varnishes with pigments or paints with addition of resin or gum. The base used is usually zinc white, because of its hard finish and fine grain. They are preferred to paints because of the brilliant glossy finish, which sheds dust and outlasts paint. They are, however, more difficult to apply, although this can be remedied by the addition of a little linseed oil.

570. Wood stains. Oil stains are very thin paints with specially ground colors, thinned with turpentine. They do not raise the

grain, are easy to make and apply, but do not penetrate. Spirit stains are solutions of dyes in oil or alcohol. They are good for penetration. Chemical reactions are sometimes made use of for the staining of wood, as in the action of iron salts on the tannin in oak and chestnut, to produce very dark colors.

Shingle stains are a mixture of benzene, creosote oil, drier, and color. The creosote acts as a preservative.

571. Fillers. These are meant to fill the pores of the wood and bring the resins to the surface, to furnish a good foundation for subsequent coats. Many of them are nothing but poor quality varnishes, and one might just as well use the varnish itself with a little thinner.

572. Sizing. This is practically a filler, used on plaster. It is usually a thin glue meant to make the surface non-porous.

573. Water-soluble paints or washes are usually used on rough work, where a cheap, quick job is desired. Whitewash or kalsomine is a suspension of slaked lime with whiting (chalk) in water, protected by flour paste or glue as protective colloid. When used inside, a preservative, like alum, should be added. Cement wash is similar to whitewash, but is superior for outside work. Calci-mine is a mixture of whiting with glue or casein. Pigments are frequently added to these water paints and the article sold under fancy names for interior decoration. They all have a good deal of glue, which protects the suspended matter, making it flow easily from the brush. They are therefore very easily applied in one coat, cover anything at all without preliminary treatment, but cannot be washed and are therefore not lasting.

574. Removing paint and varnish. To remove paint or varnish from a wood or plaster surface, mechanical methods are most often employed. But they are very tedious and the work may be done by the application of alkalies. A mixture of one part soda lye (NaOH) and three parts quicklime, applied as a paste, with a cheap fiber brush, will, in about twelve hours, soften the paint so that it may be washed off with a weak solution of washing soda.

The surface should then be neutralized with a little vinegar or dilute hydrochloric acid. On a floor, a scrubbing brush and lye will remove varnish at once.

575. Washing paint. Paint and varnish should obviously not be washed with strong alkalies, like soda, or even strong soap. If the paint is particularly dirty, it may be quickly and easily cleaned with a cotton cloth and a weak soda solution; but since this removes a thin layer of paint, if applied too often, repainting becomes necessary. This washing with soda serves fairly well as a preliminary to repainting. To remove stains of varnish or paint, turpentine or carbon tetrachloride may be used. Soda solution removes stains effectively, if in a place where soda may be applied, like the hands. Pitch or asphalt is easily removed by gasoline or kerosene. The white stains produced on highly polished wood by water or hot articles may be removed by polishing, locally, with an emulsion of linseed oil, turpentine, and alcohol.

576. Preparation of a surface for painting. Soft porous surfaces, like plaster, must first be sized with glue, holes filled in, and irregularities smoothed out. Wood surfaces must be filled, cracks and holes puttied, and the entire surface smoothed, for a good job. Metal surfaces must be cleansed of rust with a wire brush. Iron may be treated with a solution of a mixture of nitric acid and copper sulphate. This dissolves rust and plates the surface with copper in one operation. It must then be washed and dried before painting. Old paint must be removed (or at least all loose stuff) before applying new paint. Powdery white lead paint is easily incorporated with the new paint and need therefore only be dusted with a dry brush. Knots in wood must be treated with shellac before painting, especially where subject to the sun, which brings the resins to the surface.

For outside work, the first coat should consist of a thin suspension of a white lead paste in turpentine. The second coat should be a paint composed of half lead and half zinc. The third coat should

be two thirds zinc and one third lead ; or for a high gloss, a spar varnish may be substituted.

577. Inks. *Printing inks* are practically paints, made of the usual pigments, the most common of which is boneblack, ground in burned linseed oil, with manganese borate as a drier.

Drawing inks. India ink consists of lampblack, in colloidal suspension, protected by glue, gum arabic, or shellac. Fluidity may be increased in the manner usual with colloids: addition of a little alkali, such as ammonia water ; and in fact, the ink may also be removed by this means.

Marking ink. Linens sent to the laundry are often marked with silver nitrate solution. This is reduced on the cloth to metallic silver, which, like all metals, in coarse dispersion, is black (colloidal).

Writing inks. Some of these are merely solutions of dyestuffs. Ordinary red ink is a water solution of eosine. " Ink crystals " consist essentially of nigrosine, a black dye, which, with water, makes a solution frequently used as ink. None of the dyestuffs, however, make permanent writing, since, at any time, it may be removed by water.

Permanent writing fluids usually contain ferrous tannate, a colorless solute, which, on exposure to air, oxidizes to ferric tannate, an insoluble black precipitate ; this, formed in the meshes of the paper fibers, cannot easily be removed, mechanically, or by solution. The following materials are used in producing writing ink : tannic acid ; water ; ferrous sulphate ; dyestuff (usually blue) ; preservative (phenol) ; gum arabic. The purpose of the first three is, obviously, to produce the solution of ferrous tannate. Now, up to modern times, this was always exposed to air during the manufacture in order to darken it by oxidation ; but this has two disadvantages: much of the black precipitate of ferric tannate settles out, producing a muddy liquid, and at the same time this precipitation, having already taken place, cannot do so " in the meshes of the paper " ; so the writing is chiefly a surface

affair, and not permanent. Such ink would never do for fountain pens, because it is not fluid enough; therefore our modern inks are not oxidized beforehand, and since ferrous tannate is almost colorless, it becomes necessary to add a little dyestuff. This gives immediate color to the fluid, but after the writing has been exposed to the air for a few days, a good ink will turn black, due to formation of ferric tannate. Naturally, as the bottle of ink is opened again and again, it will oxidize and throw down a little ferric tannate; but if we have a protective colloid, like gum arabic, present, this process will be slowed up, and whatever ferric tannate is formed will be held in colloidal suspension. This keeps the ink properly fluid. The preservative is used to protect the organic material from the action of molds and other enzymes.

Indelible writing fluids. Many preparations, called indelible inks, may be found on the market, but practically none of them are actually indelible, unless they are thin paints and not inks. There is, however, one simple effective way to produce absolutely indelible writing on paper. A mixture of one part concentrated sulphuric acid with three parts of water may be used in a clean fountain pen; and if the writing, at first colorless, is warmed on a steam radiator, it turns a beautiful jet-black, due to dehydration of the paper and production of carbon. At the same time the paper is etched by the acid, so that, even if the carbon, insoluble in all reagents, were removed in colloidal condition, the etched writing would remain.

Sympathetic ink. Writing with a dilute solution of cobalt nitrate is almost colorless. If the paper be now warmed, the writing turns dark blue. (Cobalt ions are pink, the molecule is blue.) This is the material sometimes used in household "barometers." When the air is moist, the cobalt nitrate becomes pink; dry air takes up the water, and the salt turns blue.

A thin starch solution mixed with potassium iodide is colorless. Addition of hydrogen peroxide or chlorine water changes the solu-

tion to blue, due to liberation of iodine, which, with starch, gives the color.

In both these cases, the writing is at first invisible, and only becomes apparent on heating, or by the addition of another substance. Such fluids are called *sympathetic inks*.

SUPPLEMENTARY READING

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H. E. Howe and F. M. Turner, *Chemistry in the Home* (Charles Scribner's Sons, 1927), chap. ix (Ink).

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QUESTIONS

1. Why does a rough piece of wood decay more easily than a smooth one?

2. Why is the part of a post which is below the ground painted with creosote?

3. What are two essential constituents of a paint?

4. Why would it not be satisfactory to paint our houses with dye?

5. What is a vehicle, a thinner, a drying oil, a base?

6. What is the difference between a base and a pigment?

7. Of all the oils available, why is linseed oil selected?

8. Compare, giving advantages and disadvantages, zinc white and white lead.

9. Old paintings sometimes darken in color. To what may this be due? Hydrogen peroxide will change the dark to a light color. Explain.

10. Name a good blue pigment; green; yellow; red.

11. What are the properties of turpentine which make it so valuable as a constituent of paints?

12. What is a varnish, an enamel, an extender, a filler, a lacquer, whitewash?
13. Why is it necessary to prepare a surface so carefully before painting it?
14. How would you remove (*a*) a paint; (*b*) a varnish?
15. What is an essential difference between writing and printing ink?
16. Why is it necessary to stir paint with a wooden paddle before use?
17. Why do paints settle out on standing?
18. Mention other ways of protecting a surface besides painting.
19. Why do brick and rock structures require no paint?

OPTIONAL QUESTIONS

20. Describe how one might produce a paint which would not settle.
21. There are a number of water paints containing glue. Explain the purpose of using glue.

CHAPTER XXXVI

CHEMISTRY IN AMERICA

578. In 1620, leather, salt, and iron were manufactured on a small scale in Virginia, and two years later these early chemical industries were extinguished by the Indian massacre. In 1623 the manufacture of salt was commenced by the Plymouth colony in New England, and within the next ten years a brick kiln and a tannery were established in the state. The first paper to be manufactured by the American colonies was in Germantown, Pa., in 1697. In 1743 the American Philosophical Society (which included science) was established by Benjamin Franklin, and in 1769 Benjamin Rush was appointed to the first professorship of chemistry in America at the Philadelphia Medical College. In 1793 John Harrison, of Philadelphia, established the manufacture of sulphuric acid. The following year chemistry in America received a fresh impetus by the arrival from England of Joseph Priestley, the discoverer of oxygen.

Some of the more outstanding achievements in the development of American chemistry may now be summarized in chronological order:

- 1801, Robert Hare invents the oxy-hydrogen blowpipe.
- 1802, Du Pont establishes his powder works in Wilmington, Del.
- 1806, David Melville introduces coal gas for lighting in Newport, R. I.
- 1811, Joseph Hawkins, of Philadelphia, introduces the manufacture of carbonated water.
- 1812, anthracite coal is being introduced in the industries.
- 1816, gas lighting on a large scale is undertaken in Philadelphia.
- 1818, Benjamin Silliman establishes the *American Journal of Science and Arts*.

- 1819, petroleum is discovered in boring salt wells in Ohio.
- 1819, the first canned goods are put up in New York.
- 1825, Isaac Babbitt, of Taunton, Mass., invents Babbitt metal.
- 1829, John M. Revere, of New York, perfects a process for galvanizing iron.
- 1831, Samuel Guthrie discovers chloroform.
- 1832, use of tanbark in the manufacture of white lead is introduced.
- 1837, first beet sugar is produced in America in Northampton, Mass.
- 1839, Charles Goodyear discovers the process for vulcanizing rubber.
- 1844, W. A. Burt discovers the Lake Superior iron-ore deposits.
- 1851, William Kelly invents a process for converting pig iron into malleable iron and steel.
- 1856, Bessemer steel is first made in Phillipsburg, N. J.
- 1863, the National Academy of Sciences is incorporated by Congress.
- 1867, B. C. Tilghman invents the sulphite pulp process for the manufacture of paper.
- 1872, D. O. Saylor begins manufacture of Portland cement in the United States.
- 1874, the American Association for the Advancement of Sciences is founded.
- 1876, the American Chemical Society is founded.
- 1879, Remsen discovers saccharin.
- 1879, *Journal of the American Chemical Society* is founded.
- 1886, C. M. Hall prepares aluminum electrolytically.
- 1888, T. L. Wilson introduces the electric process for making calcium carbide.
- 1890, H. Y. Castner introduces the electrolytic process for making sodium.
- 1891, E. G. Acheson produces carborundum.
- 1896, E. G. Acheson produces artificial graphite.
- 1902, the beginning of a large production of gasoline for automobiles.

1905, L. H. Baekeland discovers "bakelite."

1907, H. P. Cady discovers helium in natural gas in Kansas.

1910, Irving Langmuir introduces his gas-filled lamp.

1910, opening of the Chemists Building in New York City.

1912, W. M. Burton introduces the process for cracking petroleum.

1913, the Standard Chemical Company begins the commercial production of radium.

Since 1914 the chemical industries in this country have grown and developed to a remarkable extent; this has shown itself particularly in the synthetic production of dyes and drugs. Side by side with the growth of the chemical industries, the study of chemistry has received more and more attention at the universities.

579. Since 1914 Millikan has measured the electron; Langmuir has developed the electron theory; Hopkins has isolated a new element, "illinium"; Banting has obtained an active extract ("insulin"), of value in the treatment of diabetes; a highly improved glass (pyrex) has been developed; Cottrell has devised a method for eliminating smoke; the potash industry has been developed in the West; methods for the extraction of iodine from kelp have been devised; large nitrogen fixation plants have been built; the cellulose industry (rayon) has assumed enormous proportions; and much valuable work on the chemistry of foods (vitamins, etc.) has been accomplished.

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- E. F. Smith, "The History of Chemistry in America, with Special Reference to Yale," *Journal of Industrial and Engineering Chemistry*, vol. XV (1923), p. 464.
- C. A. Browne, "Chronological Table of Some Leading Events in the History of Industrial Chemistry in America," *Journal of Industrial and Engineering Chemistry*, vol. XVIII (1926), p. 884.

REVIEW

LABORATORY PREPARATIONS

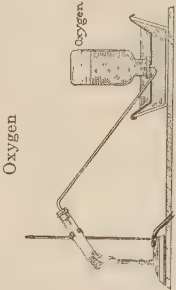
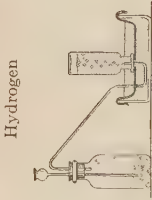
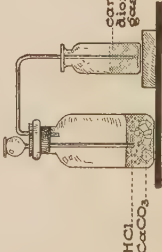
DEFINITIONS

EQUATIONS

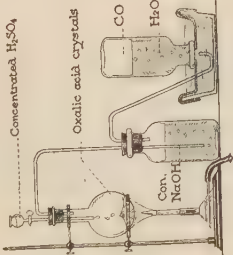
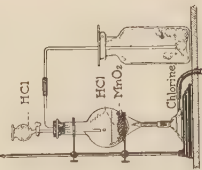
COMMON SUBSTANCES

TABLE OF SOLUBILITIES

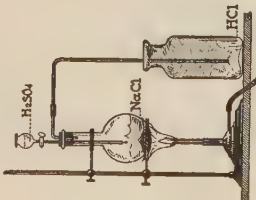
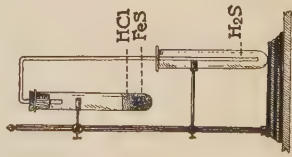
SOME COMMON LABORATORY PREPARATIONS

LABORATORY SET-UP	EQUATION FOR PREPARATION	PHYSICAL PROPERTIES	CHEMICAL PROPERTIES	USES	TEST
Oxygen 	$2 \text{KClO}_3 \xrightarrow{\text{MnO}_2} 2 \text{KCl} + 3 \text{O}_2$	Colorless, odorless, and tasteless; slightly heavier than air; slightly soluble in water.	Combines with most elements forming oxides. Supports combustion and life.	Oxy-hydrogen blowpipe; rescue parties in mines; aviators.	A glowing splinter bursts into flame in this gas.
Hydrogen 	$\text{Zn} + \text{H}_2\text{SO}_4 \rightarrow \text{H}_2 + \text{ZnSO}_4$	Colorless, odorless, and tasteless; insoluble in water and lightest of all substances.	Reducing agent. Burns with blue flame.	Hydrogenation of oils; inflation of airships; oxyhydrogen blowpipe.	Forms water with oxygen.
Carbon dioxide 	$\text{CaCO}_3 + 2 \text{HCl} \rightarrow \text{H}_2\text{O} + \text{CO}_2 + \text{CaCl}_2$	Colorless, odorless, soluble in water, heavier than air.	Combines with bases to form carbonates.	Beverages; raising dough; fire extinguisher.	Turns limewater milky.

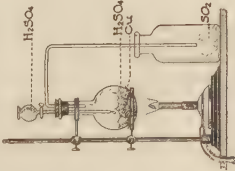
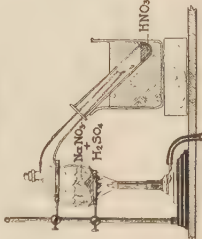
SOME COMMON LABORATORY PREPARATIONS—Continued

LABORATORY SET-UP	EQUATION FOR PREPARATION	PHYSICAL PROPERTIES	CHEMICAL PROPERTIES	USES	TEST
Carbon monoxide 	$\text{H}_2\text{SO}_4 + \text{H}_2\text{C}_2\text{O}_4 \rightarrow \text{H}_2\text{O} + \text{CO} + \text{CO}_2$	Colorless, odorless, insoluble, lighter than air.	Reducing agent. Burns with a blue flame.	Blast furnace; water gas; producer gas.	Burns to form CO_2.
Chlorine 	$\text{MnO}_2 + 4 \text{HCl} \rightarrow 2 \text{H}_2\text{O} + \text{Cl}_2 + \text{MnCl}_2$	Greenish-yellow, choking odor, soluble in water, heavier than air.	Combines with most elements forming chlorides; supports combustion; bleaches.	Bleaching; disinfecting; water purification.	Color.

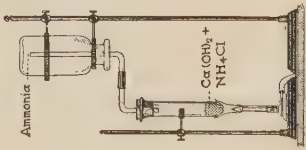
SOME COMMON LABORATORY PREPARATIONS — Continued

LABORATORY SET-UP	EQUATION FOR PREPARATION	PHYSICAL PROPERTIES	CHEMICAL PROPERTIES	USES	TEST
Hydrogen chloride 	$2 \text{NaCl} + \text{H}_2\text{SO}_4 \rightarrow 2 \text{HCl} + \text{Na}_2\text{SO}_4$	Colorless, sharp odor, very soluble.	Forms strong acid in water. Chlorides.	Pickling metals, soldering, making glue, etc.	Fumes with NH_3; water solution forms white precipitate with AgNO_3, and turns blue litmus red.
Hydrogen sulphide 	$\text{FeS} + 2 \text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2\text{S}$	Colorless, odor of rotten eggs, soluble, heavier than air.	Combines with metals; burns to form sulphur or sulphur dioxide; a reducing agent; a weak acid.	To detect metal ions in analytical chemistry.	Blackens lead acetate paper.

SOME COMMON LABORATORY PREPARATIONS — Continued

LABORATORY SET-UP	EQUATION FOR PREPARATION	PHYSICAL PROPERTIES	CHEMICAL PROPERTIES	USES	TEST
Sulphur dioxide 	$\text{Cu} + 2 \text{H}_2\text{SO}_4 \rightarrow 2 \text{H}_2\text{O} + \text{SO}_2 + \text{CuSO}_4$	Colorless, choking odor, soluble, heavier than air.	Reducing agent, acid anhydride.	Bleaching; disinfecting; manufacture of sulphuric acid.	Oxidize and test for a sulphate with BaCl_2.
Nitric acid 	$\text{NaNO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{HNO}_3 + \text{NaHSO}_4$	Heavier than water; slightly yellow.	Oxidizing agent.	Explosives; making fertilizers.	Ferrous sulphate + conc. H_2SO_4 give brown ring.

SOME COMMON LABORATORY PREPARATIONS — Continued

LABORATORY SET-UP	EQUATION FOR PREPARATION	PHYSICAL PROPERTIES	CHEMICAL PROPERTIES	USES	TESTS
Ammonia 	$\begin{aligned} &(\text{NH}_4)_2\text{SO}_4 + \\ &\text{Ca}(\text{OH})_2 \rightarrow 2\text{NH}_3 \\ &+ 2\text{H}_2\text{O} + \text{CaSO}_4 \end{aligned}$	Colorless, characteristic odor, very soluble in water, lighter than air.	Basic anhydride.	Refrigeration ; making fertilizers.	Smell.

REVIEW

DEFINITIONS

For more exact definitions, see text.

Acid. A substance containing positive hydrogen and whose water solution gives hydrogen ions.

Examples: HCl , H_2SO_4 , HNO_3 .

Alkali. A substance whose solution turns red litmus blue.

Examples: NaOH , Na_2CO_3 .

Allotropic forms. Different physical forms of the same element.

Example: graphite, diamond, and amorphous carbon are allotropic forms of carbon.

Alloy. A solid solution.

Examples: brass, bronze, coins.

Amalgam. An alloy containing mercury.

Example: silver amalgam.

Amorphous. Without crystalline form.

Examples: plastic sulphur; glue.

Analysis. Decomposition of a substance.

Example: $2 \text{HgO} \longrightarrow \text{O}_2 + 2 \text{Hg}$.

Anhydride. A substance which, on addition of water, becomes (a) an acid or (b) a base.

Examples: (a) $\text{SO}_3 + \text{H}_2\text{O} \longrightarrow \text{H}_2\text{SO}_4$

(b) $\text{CaO} + \text{H}_2\text{O} \longrightarrow \text{Ca(OH)}_2$

Anhydrous. Without water.

Example: Washing soda after losing water of hydration.

Anode. The positive electrode.

Example: In the electrolysis of water, it is the pole at which oxygen is set free.

Atom. The smallest particle of an element taking part in a chemical change.

Example: The smallest particle of hydrogen that can combine with chlorine is an atom.

Atomic weight. The number of times one atom of an element is as heavy as one atom of hydrogen.

Example: When we say that the atomic weight of oxygen is 16, we mean that one atom of oxygen is 16 times as heavy as one atom of hydrogen.

Base. A substance containing negative OH, and whose water solution gives hydroxyl ions.

Examples: KOH, NH_4OH , $\text{Ca}(\text{OH})_2$.

Carbohydrate. A name given to a group of compounds including the sugars, starch, and cellulose.

Catalytic agent. A substance which changes the speed of a reaction without itself undergoing a permanent change.

Example: colloidal platinum in the manufacture of sulphuric acid.

Cathode. The negative electrode.

Example: In the electrolysis of water, it is the pole at which hydrogen is set free.

Chemical change. The production of a substance with properties totally different from the original. Chemical changes involve a transfer of one or more electrons.

Colloid. A substance whose particles are intermediate in size between coarse suspensions and molecules.

Example: Glue (a colloid) has particles smaller than clay (a suspension) and yet larger than the molecules of salt dissolved in water.

Combustion. A rapid chemical action producing light and heat.

Examples: burning of paper; burning of sodium in chlorine.

Compound. When two or more elements combine in definite proportions by weight, a compound is formed.

Example: 8 grams of oxygen combine with 1 gram of hydrogen to form 9 grams of water. Water is a compound.

Decomposition. Breaking up of a compound into simpler substances.

Example: electrolysis of water.

Deliquescent substance. One which takes up water from the air and gets wet.

Example: calcium chloride.

Destructive distillation. The process of decomposing by heat a complex organic substance, in the absence of air.

Example: The destructive distillation of coal yields coke, ammonia, coal tar, and coal gas.

Distillation. The separation of a liquid from a mixture by boiling and condensing the vapor.

Example: In the distillation of impure water, the pure water is separated from its impurities.

Efflorescent substance. The loss of water of hydration by certain crystalline substances on exposure to air.

Example: washing soda.

Electrolysis. The decomposition of a substance by electricity.

Example: the electrolysis of brine to yield sodium hydroxide.

Electrolyte. A substance whose solution conducts the electric current and is decomposed by it.

Examples: sulphuric acid; sodium hydroxide; potassium nitrate.

Element. A substance which, in a chemical reaction, has not been decomposed into anything simpler than itself.

Example: Iron is called an element because, in a chemical reaction, it does not decompose into anything simpler than iron.

Equilibrium. The point in a reversible reaction at which the rate of chemical change proceeding in one direction is exactly equal to that in the other direction.

Example: Heating calcium carbonate in a closed vessel gives some calcium oxide and some carbon dioxide, but there is still some calcium carbonate left.



Flux. A substance used in metallurgical processes to remove impurities by forming a liquid with them.

Example: the use of limestone (the flux) in removing the silica from iron ore, forming *slag* (a liquid).

Hydrate. A crystalline substance containing a definite amount of water, in combination.

Example: blue vitriol, $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$.

Hydrocarbon. A compound containing hydrogen and carbon.

Examples: benzene, C_6H_6 and acetylene, C_2H_2 .

Hydrolysis. The combination of water with a substance producing new substances. It is usually the reverse of neutralization.

Example: $\text{Na}_2\text{CO}_3 + 2 \text{H}_2\text{O} \rightleftharpoons 2 \text{NaOH} + \text{H}_2\text{CO}_3$.

Ion. The loss or gain of one or more electrons by an atom or group of atoms (radical) produces an ion.

Examples: Na^+ , Cl^- , SO_4^{--} , NH_4^+ .

Isotopes. An element may consist of two or more kinds of atoms, which differ slightly in atomic weight, but have the same chemical properties. These different forms of the same element are called isotopes.

Example: Ordinary chlorine gas is a mixture of two isotopes, with atomic weights of 35 and 37.

Kindling temperature. The lowest temperature at which a substance catches fire.

Example: The kindling temperature of phosphorus is 37°C .

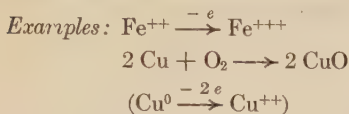
Molecular weight. The sum of the weights of all the atoms used in building up the molecule.

Examples: $\text{O}_2 = 16 + 16 = 32$; $\text{H}_2\text{SO}_4 = 2 + 32 + 64 = 98$.

Neutralization. The combination of the hydrogen ions of an acid and the hydroxyl ions of a base to form undissociated water.

Examples: $\text{Na}^+ + \text{OH}^- + \text{H}^+ + \text{Cl}^- \longrightarrow \text{H}_2\text{O} + \text{Na}^+ + \text{Cl}^-$
 $\text{Ca}^{++} + 2 \text{OH}^- + 2 \text{H}^+ + \text{CO}_3^{--} \longrightarrow \text{CaCO}_3 \downarrow + 2 \text{H}_2\text{O}$

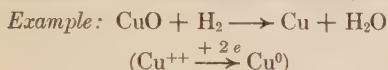
Oxidation involves a change from a lower to a higher positive valence.



Precipitate. An insoluble substance formed by the reaction of two clear solutions.



Reduction. A change from a higher to a lower positive valence.

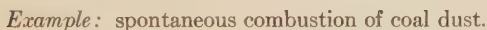


Salt. A compound consisting of a positive ion other than hydrogen and a negative ion other than hydroxyl. (Salts may be formed in four ways; see text.)



Saturated solution. When an excess of a solute is added to a solvent, that portion of the substance which goes into solution forms a saturated solution.

Spontaneous combustion. The starting of a fire without the external application of heat.



Sublimation. The change from a solid to a vapor without the apparent formation of a liquid.



Synthesis. The building up process.

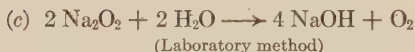
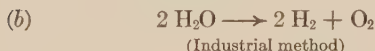
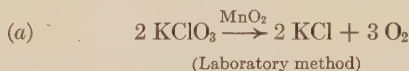


Valence. The number which indicates how many electrons an element or radical has gained or lost in chemical action.

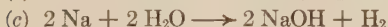
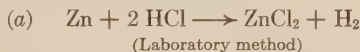


EQUATIONS

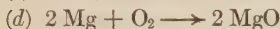
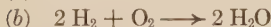
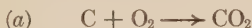
1. Preparation of oxygen :



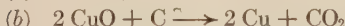
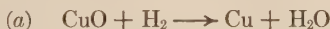
2. Preparation of hydrogen :



3. Combustion or oxidation :



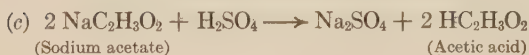
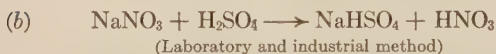
4. Reduction :



5. Chemical action in solution :



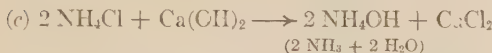
6. Preparation of acids :



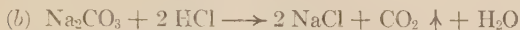
7. Preparation of bases :



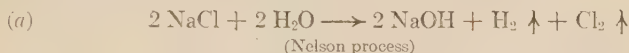
(Laboratory and industrial method)



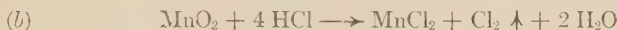
8. Preparation of a salt by different methods :



9. Halogens :



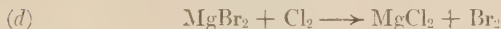
(Nelson process)



(Laboratory method)



(Laboratory preparation of a halogen. $\text{X} = \text{Cl}, \text{Br}, \text{or I}$)



(Industrial method)

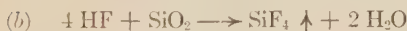


(Replacement of iodine)

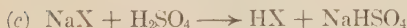
10. Halogen acids :



(Industrial process)

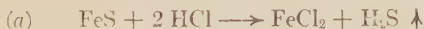


(Etching of glass)



(Laboratory preparation of halogen acids. $\text{X} = \text{Cl}, \text{Br}, \text{or I}$)

11. Compounds of sulphur :



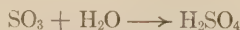
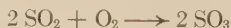
(Laboratory method)



(Test for H_2S)



(Laboratory method)



("Contact process" for sulphuric acid)

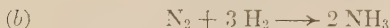


(Test for a sulphate)

12. Compounds of nitrogen :



(Laboratory method)



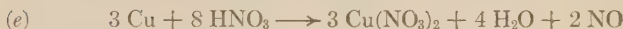
(Haber process)



(Laboratory method)



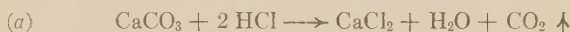
(Industrial process)



13. Chemistry in agriculture :



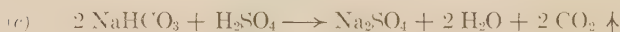
14. Oxides of carbon :



(Laboratory method)



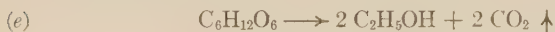
(Test for CO_2)



(A fire extinguisher)

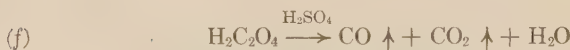


(Baking powder)



(glucose) (alcohol)

(Fermentation)



(Oxalic acid) (Dehydrating agent)

(Laboratory method for preparing CO)



(Reduction)

- (b) $\text{CaO} + \text{H}_2\text{O} \longrightarrow \text{Ca(OH)}_2$
(Slacking of lime)
- (c) $\text{Ca(OH)}_2 + \text{CO}_2 \longrightarrow \text{CaCO}_3 \downarrow + \text{H}_2\text{O}$
("Setting")
- (d) $\text{CaCO}_3 + \text{H}_2\text{O} + \text{CO}_2 \longrightarrow \text{Ca(HCO}_3)_2$
(Temporary hard water)
- (e) $\text{Ca(HCO}_3)_2 \longrightarrow \text{CaCO}_3 \downarrow + \text{H}_2\text{O} + \text{CO}_2$
(Softening of temporary hard water)
- (f) $\text{CaCO}_3 + \text{SiO}_2 \longrightarrow \text{CaSiO}_3 + \text{CO}_2$
(Formation of slag)
- (g) $2[\text{CaSO}_4 \cdot 2 \text{H}_2\text{O}] \longrightarrow (\text{CaSO}_4)_2 \cdot \text{H}_2\text{O} + 3 \text{H}_2\text{O}$
(Manufacture of plaster of Paris)
- (h) $\text{Ca(OH)}_2 + \text{Cl}_2 \longrightarrow \text{Ca(OCl)Cl} + \text{H}_2\text{O}$
(Bleaching powder)
- (i) $\text{Ca(OCl)Cl} + \text{H}_2\text{SO}_4 \longrightarrow \text{CaSO}_4 + \text{H}_2\text{O} + \text{Cl}_2 \uparrow$
(Bleaching)

20. Electricity in chemistry :

- (a) $2 \text{NaOH} \longrightarrow 2 \text{Na} + \text{H}_2 + \text{O}_2$
(Manufacture of sodium)
- (b) $2 \text{NaCl} + 2 \text{H}_2\text{O} \longrightarrow 2 \text{NaOH} + \text{H}_2 + \text{Cl}_2$
(Nelson process)
- (c) $2 \text{Al}_2\text{O}_3 \longrightarrow 4 \text{Al} + 3 \text{O}_2$
(Hall process)
- (d) $\text{CaO} + 3 \text{C} \longrightarrow \text{CaC}_2 + \text{CO} \uparrow$
(Calcium carbide)
- (e) $\text{CaC}_2 + 2 \text{H}_2\text{O} \longrightarrow \text{Ca(OH)}_2 + \text{C}_2\text{H}_2$
(Acetylene)
- (f) $\text{SiO}_2 + 3 \text{C} \longrightarrow \text{SiC} + 2 \text{CO} \uparrow$
(Carborundum)

21. Carbon compounds. See text.

TABLE OF COMMON SUBSTANCES

Alum	Potassium aluminum sulphate, $\text{KAl}(\text{SO}_4)_2 \cdot 12 \text{H}_2\text{O}$.
Ammonia water	Ammonium hydroxide solution, NH_4OH .
Aqua fortis	Nitric acid, HNO_3 .
Aqua regia	A mixture of hydrochloric acid and nitric acid.
Baking soda	Sodium bicarbonate, NaHCO_3 .
Bleaching powder	Calcium chloride-hypochlorite, $\text{CaCl}(\text{OCl})$.
Blue vitriol	Crystalline copper sulphate, $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$.
Boracic acid	Boric acid, H_3BO_3 .
Calomel	Mercurous chloride, HgCl .
Carbolic acid	Phenol, $\text{C}_6\text{H}_5\text{OH}$.
Caustic	(lime, soda, potash). The hydroxide of the given alkali, or alkaline earth metal.
Chalk	A mineral, impure calcium carbonate, CaCO_3 .
Chloride of lime	A term, frequently and incorrectly applied to bleaching powder.
Choke damp	Carbon dioxide.
Chrome yellow	Lead chromate, PbCrO_4 .
Cinnabar	A mineral, mercuric sulphide, HgS .
Copperas	Ferrous sulphate, $\text{FeSO}_4 \cdot 7 \text{H}_2\text{O}$.
Corrosive sublimate	Mercuric chloride, HgCl_2 .
Cream of tartar	Potassium acid tartrate, $\text{KHC}_4\text{H}_4\text{O}_6$.
Dakin's solution	A solution of sodium hypochlorite.
Epsom salts	Magnesium sulphate, MgSO_4 .
Fire damp	Methane, CH_4 .
Glauber's salt	Sodium sulphate, $\text{Na}_2\text{SO}_4 \cdot 10 \text{H}_2\text{O}$.
Laughing gas	Nitrous oxide, N_2O .
Limewater	A solution of calcium hydroxide, $\text{Ca}(\text{OH})_2$.
Lunar caustic	Silver nitrate, AgNO_3 .
Marble	A native form of calcium carbonate, CaCO_3 .
Marsh gas	Methane, CH_4 .
Muriatic acid	Hydrochloric acid, HCl .
Niter	Potassium nitrate, KNO_3 .
Oil of vitriol	Concentrated sulphuric acid, H_2SO_4 .
Prussic acid	Hydrocyanic acid, HCN .
Pyrene	Name for carbon tetrachloride, CCl_4 , when used in fire extinguishers.
Quicklime	Calcium oxide, CaO .
Rochelle salt	Potassium sodium tartrate, $\text{KNaC}_4\text{H}_4\text{O}_6$.
Rock salt	Sodium chloride, NaCl .
Sal ammoniac	Ammonium chloride, NH_4Cl .
Sal soda	Sodium carbonate, Na_2CO_3 .

TABLE OF COMMON SUBSTANCES—*Continued*

Saleratus	Sodium or potassium bicarbonate, NaHCO_3 or KHCO_3
Salt, common	Sodium chloride, NaCl .
Saltpeter	Potassium nitrate, KNO_3 .
Saltpeter, Chile	Sodium nitrate, NaNO_3 .
Slaked lime	Calcium hydroxide, Ca(OH)_2 .
Spirits of ammonia	A solution of ammonium hydroxide.
Tartar emetic	Potassium antimonyl tartrate, $\text{K(SbO)C}_4\text{H}_4\text{O}_6$.
Vitriol, oil of	Concentrated sulphuric acid.
Water glass	Sodium silicate, Na_2SiO_3 .
Whiting	Chalk, calcium carbonate, CaCO_3 .

TABLE OF SOLUBILITIES IN WATER

	ACETATE	BROMIDE	CARBONATE	CHLORIDE	FERRICYANIDE	FERROCYANIDE	HYDROXIDE	IODIDE	NITRATE	OXALATE	OXIDE	PHOSPHATE	SILICATE	SULPHATE	SULPHIDE
Al	S	S	—	S	—	×	I	S	S	I	I	I	—	S	I
NH_4	S	S	S	S	S	S	S	S	S	S	—	S	—	S	S
Ba	S	S	I	S	×	×	S	S	S	I	S	I	I	I	S
Bi	S	×	I	×	—	×	I	I	S	I	I	I	—	S	I
Ca	S	S	I	S	S	S	×	S	S	I	×	I	I	I	S
Cr	S	S	—	S	—	—	I	S	S	×	I	I	I	S	I
Cu	S	S	I	S	I	I	I	S	S	I	I	I	I	S	I
H	S	S	—	S	S	S	S	S	S	S	S	S	S	S	×
Fe^{++}	S	S	I	S	I	I	I	S	S	I	I	I	I	S	I
Fe^{+++}	S	S	—	S	S	I	I	S	S	I	I	I	I	S	I
Pb	S	I	I	I	×	I	I	×	S	I	I	I	I	I	I
Mg	S	S	I	S	S	S	I	S	S	I	I	I	I	S	I
Hg^+	×	I	I	I	—	—	—	I	S	I	I	I	—	×	I
Hg^{++}	S	S	I	S	I	I	—	I	S	I	I	I	—	S	I
K	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S
Ag	×	I	I	I	I	I	—	I	S	I	I	I	—	×	I
Na	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S
Zn	S	S	I	S	I	I	I	S	S	I	I	I	I	S	I

S Soluble in water
 I Insoluble
 × Slightly soluble

APPENDIX

TABLE OF NUTRIENTS AND FUEL VALUES OF FOODS

FOOD	PROTEIN PER CENT	FAT PER CENT	CARBO- HYDRATE PER CENT	FUEL VALUE CALORIES PER POUND
Almonds	21.0	54.9	17.3	2940
Apples	.4	.5	14.2	285
Apricots	1.1	—	13.4	263
Asparagus, fresh	1.8	.2	3.3	100
Bacon, smoked	10.5	64.8	—	2840
Bananas	1.3	.6	22.0	447
Beans,				
Lima, dried	18.1	1.5	65.9	1586
Lima, fresh	7.1	.7	22.0	557
string, fresh	2.3	.3	7.4	184
baked, canned	6.9	2.5	19.6	583
Beef, brisket, medium fat	15.8	28.5	—	1449
chuck, average	19.2	15.4	—	978
corned, average	15.6	26.2	—	1353
cross ribs, average	15.9	28.2	—	1440
dried, salted, and smoked	30.0	6.5	.4	817
liver	20.4	4.5	1.7	584
porterhouse steak	21.9	20.4	—	1230
ribs, lean	19.6	12.0	—	845
ribs, fat	15.0	35.6	—	1721
sirloin steak	18.9	18.5	—	1099
sweetbreads	16.8	12.1	—	799
tenderloin	16.2	24.4	—	1290
tongue	18.9	9.2	—	717
Beet, cooked	2.3	.1	7.4	180
Blackberries	1.3	1.0	10.9	262
Bluefish	19.4	1.2	—	402
Brazil nuts	17.	66.8	7.	3162
Bread, Boston brown	6.0	6.3	54.0	1345
graham	8.9	1.8	52.1	1189
rolls	9.0	3.0	54.2	1268
toasted	11.5	1.6	61.2	1385
white	9.1	1.6	53.3	1199
whole wheat	9.7	.9	49.7	1113
Butter	1.0	85.0	—	3491
Buttermilk	3.0	.5	4.8	162
Cabbage	1.6	.3	5.6	143

TABLE OF NUTRIENTS AND FUEL VALUES OF FOODS—*Continued*

Food	PROTEIN PER CENT	FAT PER CENT	CARBO- HYDRATE PER CENT	FUEL VALUE CALORIES PER POUND
Carrots, fresh	1.1	.4	9.3	204
Cauliflower	1.8	.5	4.7	139
Celery	1.1	.1	3.3	84
Cheese, American pale	28.8	35.9	.3	1990
cheddar	27.7	36.8	4.1	2080
cottage	20.9	1.0	4.3	499
full cream	25.9	33.7	2.4	1890
Roquefort	22.6	29.5	1.8	1645
Swiss	27.6	34.9	1.3	1945
Cherries, fresh	1.0	.8	16.7	354
Chestnuts, fresh	6.2	5.4	42.1	1098
Chicken, broilers	21.5	2.5	—	493
Chocolate	12.9	48.7	30.3	2768
Cocoa	21.6	28.9	37.7	2258
Cod, dressed	11.1	.2	—	209
Corn, green, canned	2.8	1.2	19.0	455
sweet, fresh	3.1	1.1	19.7	459
Crackers, graham	10.0	9.4	73.8	1905
soda	9.8	9.1	73.1	1875
Cranberries	.4	.6	9.9	212
Cream	2.5	18.5	4.5	883
Cucumbers	.8	.2	3.1	79
Dates, dried	2.1	2.8	78.4	1575
Doughnuts	6.7	21.0	53.1	1941
Eggs	13.4	10.5	—	672
Farina	11.0	1.4	76.3	1640
Figs, dried	4.3	.3	74.2	1437
Flounder	14.2	0.6	—	290
Gelatin	91.4	.1	—	1660
Grapefruit	.6	.1	12.2	235
Grapes	1.3	1.6	19.2	437
Haddock	17.2	.3	—	324
Halibut steaks	18.6	5.2	—	550
Ham, fresh	15.3	28.9	—	1458
smoked	19.8	20.8	—	1209
Herring, whole	19.5	7.1	—	644
Honey	.4	—	81.2	1481
Huckleberries	.6	.6	16.6	336
Lamb, breast	19.1	23.6	—	1311
chops, broiled	21.7	29.9	—	1614
leg, roast	19.7	12.7	—	876

TABLE OF NUTRIENTS AND FUEL VALUES OF FOODS—*Continued*

FOOD	PROTEIN PER CENT	FAT PER CENT	CARBO- HYDRATE PER CENT	FUEL VALUE CALORIES PER POUND
Lard, refined	—	100.0	—	4080
Lemons	1.0	.7	8.5	201
Lettuce	1.2	.3	2.9	87
Liver, beef	20.4	4.5	1.7	584
veal	19.0	5.3	—	562
Lobster, whole	16.4	1.8	.4	379
Mackerel	18.7	7.1	—	629
Marmalade, orange	.6	.1	84.5	1548
Milk, condensed, sweetened	8.8	8.3	54.1	1480
skimmed	3.4	.3	5.1	167
whole	3.3	4.0	5.0	314
Mushrooms	3.5	.4	6.8	204
Mutton, leg	19.8	12.4	—	863
Oatmeal	16.1	7.2	67.5	1811
Olives, green	1.1	27.6	11.6	1357
Onions, fresh	1.6	.3	9.9	220
Oranges	.8	.2	11.6	233
Oysters in shell	1.2	.2	.7	43
Parsnips	1.6	.5	13.5	294
Peaches, canned	.7	.1	10.8	213
fresh	.7	.1	9.4	188
Peanuts	25.8	38.6	24.4	2490
Pears, fresh	.6	.5	14.1	288
Peas, canned	3.6	.2	9.8	252
green	7.0	.5	16.9	454
Pea soup, canned	3.6	.7	7.6	232
Pies, apple	3.1	9.8	42.8	1233
custard	4.2	6.3	26.1	806
lemon	3.6	10.1	37.4	1156
mince	5.8	12.3	38.1	1300
Pineapples, fresh	.4	.3	9.7	196
canned	.4	.7	36.4	695
Pistachios, shelled	22.3	54.0	16.3	2900
Plums	1.0	—	20.1	383
Pork, chops, medium	16.6	30.1	—	1530
sausage	13.0	44.2	1.1	2030
tenderloin	18.9	13.0	—	875
Potato chips	6.8	39.8	46.7	2598
Potatoes, white, raw	2.2	.1	18.4	378
sweet, raw	1.8	.7	27.4	558
Prunes, dried	2.1	—	73.3	1368

TABLE OF NUTRIENTS AND FUEL VALUES OF FOODS—*Continued*

FOOD	PROTEIN PER CENT	FAT PER CENT	CARBO- HYDRATE PER CENT	FUEL VALUE CALORIES PER POUND
Pumpkins	1.0	.1	5.2	117
Radishes	1.3	.1	5.8	133
Raisins	2.6	3.3	76.1	1562
Raspberries, red	1.0	—	12.6	247
Rhubarb	.6	.7	3.6	105
Rice	8.0	.3	79.0	1591
Salmon, whole	22.0	12.8	—	923
Sausage, Bologna	18.7	17.6	.3	1061
Shad, whole	18.8	9.5	—	727
roe	20.9	3.8	2.6	582
Shredded wheat	10.5	1.4	77.9	1660
Spinach, fresh	2.1	.3	3.2	109
Squash	1.4	.5	9.0	209
Strawberries	1.0	.6	7.4	169
Sugar	—	—	100.0	1815
Tomatoes, fresh	.9	.4	3.9	104
canned	1.2	.2	4.0	103
Tuna (tunny fish)	26.6	11.4	—	946
Turkey	21.1	22.9	—	1320
Turnips	1.3	.2	8.1	178
Veal, breast	20.3	11.0	—	817
cutlet	20.3	7.7	—	683
Vegetable soup, canned	2.9	—	.5	62
Walnuts, California or English	18.4	64.4	13.0	3199
black	27.6	56.3	11.7	3011
Watermelons	.4	.2	6.7	136
Whitefish	22.9	6.5	—	680
Zwieback	9.8	9.9	73.5	1915

FOODS THAT CONTAIN VITAMINS A, B, AND C

× — indicates small amount of vitamins
 ×× — a fair amount
 ××× — a relatively large amount
 ×××× — an abundance
 ××××× — an exceptionally large amount

FOOD	VITAMIN A	VITAMIN B	VITAMIN C
Almonds	×	×××	
Apples, raw	×	××	××
Asparagus		××××	
Bacon	×	××	
Bananas	×	×	××
Beans, kidney		×××	
Beans, navy		×××	
Beef	×	×	×
Beets, roots		×	
Brazil nuts		×××	
Butter	××××		
Cabbage	××××	××××	×××××
Cantaloupe	××	××	
Carrots	××××	××××	××
Cauliflower	×	××	×
Celery		××	
Cheese, whole milk	×××	×	
Chestnuts		×××	
Coconuts	××	×××	
Cod-liver	×××××		
Corn, white		××××	
Corn, yellow	××	××××	
Cranberries			×
Cucumbers		×	
Eggs, white	×		
Eggs, yolk	××××	×××	×××
Fish	×	×	
Fish (roe)	××××	××	
Grapefruit		×××	×××××
Ham	×	×××	×
Ice cream	×××	×××	×
Lemon juice		×××	×××××
Lettuce	××××	×××	××××
Milk, whole	×××	×××	×

FOODS THAT CONTAIN VITAMINS A, B, AND C — *Continued*

Food	VITAMIN A	VITAMIN B	VITAMIN C
Milk, condensed	×	×××	×
Milk, evaporated	×××	×××	
Milk, powdered	×××	×××	
Mushrooms		×××	
Mutton	×	×××	
Oats	××	××××	
Olive oil			
Onions		×	×××
Orange juice	××	×××	××××
Oysters			×××
Parsnips	××	×××	
Peaches	×	××	×××
Peanuts	×	×××	
Peas	×××	×××	×××
Pecans		×××	
Pineapples	××	×	×××
Pork	×	×××	×
Potatoes, sweet	×××	×××	×××
Potatoes, white	×	×××	×××
Prunes	×××	×××	
Radishes	×	××	
Raisins		××	×
Spinach	××××	×××	××××
Squash	××		
Strawberries		××	×
Sweetbread	××	××	
Tomatoes	××××	×××	××××
Turnip	×	×××	×××
Walnuts, English		×××	
Walnuts, black		×××	
Wheat, whole	×	××××	
Yeast		××××	

MINERAL PRODUCTS OF THE UNITED STATES

* MINERAL	CHIEF STATES
Aluminum	N. Y., N. C., Tenn.
Asbestos	Md., Cal., Ga., Ariz.
Asphalt	Cal., Tex., Ill., Ky., Utah, Okla.
Bauxite	Ark., Ga., Tenn., Ala.
Borates	Cal., Nev.
Bromine	Mich., W. Va., Ohio
Cement	Pa., Cal., Ind., Mich.
Clay products	Ohio, Pa., N. J., Ill.
Coal:	
Bituminous	Pa., W. Va., Ill., Ky.
Anthracite	Pa.
Coke	Pa., Ind., Ohio, Ill., Ala.
Copper	Ariz., Mont., Utah, Mich.
Emery	Va., N. Y.
Feldspar (crude)	N. C., Me., N. H., N. Y.
Gold	Cal., Col., S. D., Alaska
Graphite	Ala., Tex., R. I., Mich.
Gypsum	N. Y., Iowa, Ohio, Mich.
Iron ore	Minn., Mich., Ala., N. Y.
Iron, pig	Pa., Ohio, Ill., Ind., Ala.
Lead	Mo., Idaho, Utah, Okla.
Lime	Ohio, Pa., Mass., Mo.
Magnesium	N. Y., Mich.
Magnesium sulphate	Mich., Wash., Cal.
Mercury	Cal., Tex., Nev., Ore.
Mica	N. C., N. H., N. M., Va.
Natural gas	W. Va., Pa., Okla., Cal.
Peat	Ill., N. J., Cal., Ind.
Petroleum	Okla., Cal., Tex., Wyo.
Phosphate rock	Fla., Tenn., Idaho, Ky.
Platinum and allied metals	Cal., Ore., Alaska, Utah
Potash	Cal., Md., Pa., Ind.
Pyrites	Cal., Va., N. Y., Wis.
Salt	Mich., N. Y., Ohio, Kan.
Silver	Utah, Mont., Nev., Idaho
Slate	Pa., Vt., N. Y., Me.
Sulphur	Tex., La., Nev., Utah
Talc and soapstone	N. Y., Va., Vt., Cal.
Tin	Alaska
Zinc	Okla., Kan., N. J., Mont.

GROWTH IN PRODUCTION IN THE UNITED STATES

SUBSTANCE	1900	1910	1920	1926
Coal . . .	240,000,000 tons	448,000,000 tons	577,000,000 tons	600,000,000 tons
Copper . .	270,000 tons	480,000 tons	540,000 tons	770,000 tons
Corn . . .	2½ billion bushels	3 billion bushels	3⅓ billion bushels	4½ billion bushels
Cotton . .	10,000,000 bales	11,500,000 bales	13,500,000 bales	18,000,000 bales
Iron (pig) .	13,000,000 tons	27,000,000 tons	37,000,000 tons	39,000,000 tons
Petroleum .	2,670,000,000 gallons	8,800,000,000 gallons	18,600,000,000 gallons	32,000,000,000 gallons
Wheat . .	600,000,000 bushels	635,000,000 bushels	833,000,000 bushels	832,000,000 bushels

APPENDIX

THE CHEMICAL ELEMENTS

ELEMENT	SYMBOL	VALENCE	ATOMIC NO.	ATOMIC WEIGHT	SPECIFIC GRAVITY	MELTING POINT, C.	DISCOVERER
Aluminum	Al	3	13	26.96	2.702	660	Wöhler
Antimony	Sb	3, 5	51	121.77	6.684	630.5	Valentine
Argon	A	0	18	39.91	0.0178	— 189.2	Rayleigh, Ramsay
Arsenic	As	3, 5	33	74.96	2.0 to 5.7	814.0	Magnus
Barium	Ba	2	56	137.37	3.5	850	Davy
Beryllium	Be	2	4	9.02	1.85	1280	Vauquelin
Bismuth	Bi	3, 5	83	209.0	9.80	271.0	Valentine
Boron	B	3	5	10.82	2.45	2200?	Guy-Lussac, Thenard
Bromine	Br	1	35	79.92	3.4	— 7.2	Balard
Cadmium	Cd	2	48	112.41	8.6	320.9	Stromeyer
Caesium	Cs	1	55	132.81	1.90	26.0	Bunsen, Kirchoff
Calcium	Ca	2	20	40.07	1.5	810	Davy
Carbon	C	4	6	12.00	2.25	3500	—
Cerium	Ce	3, 4	58	140.25	6.90	640	Mosander
Chlorine	Cl	1	17	35.46	0.0032	— 101.5	Scheele
Chromium	Cr	2, 3, 6	24	53.01	7.1	1615.0	Vauquelin
Cobalt	Co	2, 3	27	58.97	8.9	1480	Brandt
Copper	Cu	1, 2	29	63.57	8.92	1083.0	—
Dysprosium	Dy	3	66	162.52	?	?	Boisbaudran
Erbium	Er	3	68	167.7	4.77	1250	Mosander
Europium	Eu	3	63	152.0	—	?	Demarcay
Fluorine	F	1	9	19.0	0.0017	— 223	Moissan
Gadolinium	Gd	3	64	157.26	?	?	Marignac
Gallium	Ga	3	31	69.72	5.9	29.75	Boisbaudran
Germanium	Ge	4	32	72.38	5.36	958.5	Winkler
Gold	Au	1, 3	79	197.2	19	1063.0	—
Hafnium	Hf	4	72	178.6	?	1700	Coster, Hevesy
Helium	He	0	2	4.	0.00018	— 271.0	Ramsay
Holmium	Ho	3	67	163.4	?	?	Boisbaudran
Hydrogen	H	1	1	1.008	0.000089	— 259.0	Cavendish
Illinium	Il	3	61	—	?	?	Hopkins
Indium	In	3	49	114.8	7.3	155.0	Reich, Richter
Iodine	I	1	53	126.93	4.93	113.5	Courtois
Iridium	Ir	4	77	193.1	22.4	2350.	Tennant
Iron	Fe	2, 3	26	55.84	7.86	1535.0	—
Krypton	Kr	0	36	82.9	0.0037	— 169.0	Ramsay, Travers
Lanthanum	La	3	57	138.90	6.15	286.0	Mosander
Lead	Pb	2, 4	82	207.20	11.34	327.5	—
Lithium	Li	1	3	6.94	0.534	186	Arfredson
Lutecium	Lu	3	71	175.0	?	?	Welsbach, Urbain
Magnesium	Mg	2	12	24.32	1.74	651.0	Liebig, Bussy
Manganese	Mn	2, 3, 7	25	54.93	7.2	1260	Gahn
Masurium	Ma	2, 3, 7	43	—	?	2300	Noddaek
Mercury	Hg	1, 2	80	200.64	13.56	— 38.8	—
Molybdenum	Mo	4, 6	42	96.0	10.2	2620	Hjelm

THE CHEMICAL ELEMENTS — *Continued*

ELEMENT	SYMBOL	VALENCE	ATOMIC No.	ATOMIC WEIGHT	SPECIFIC GRAVITY	MELTING POINT, C.	DISCOVERER
Neodymium	Nd	3	60	144.27	6.9	840	Welsbach
Neon	Ne	0	10	20.2	?	- 248.67	Ramsay, Travers
Nickel	Ni	2, 3	28	58.69	8.90	1452.0	Cronstedt
Niobium		5	41	93.1	8.4	1950	Hatchett
Nitrogen	N	3, 5	7	14.01	0.810	- 209.60	Rutherford
Osmium	Os	6, 8	76	190.8	22.48	2700.0	Tepnant
Oxygen	O	2	8	16.00	0.001429	- 218	Priestley, Scheele
Palladium	Pd	2, 4	46	106.7	12.0	1549.	Wollaston
Phosphorus	P	3, 5	15	31.02	1.82 to 2.25	44.2	Brandt
Platinum	Pt	2, 4	78	195.23	21.45	1755.0	Wood
Polonium	Po	2, 4, 6	84	210.0	?	?	Curie
Potassium	K	1	19	39.1	0.86	62.3	Davy
Praseodymium	Pr	2	59	140.92	6.5	940	Welsbach
Radium	Ra	2	88	225.95	5 +	960	The Curies, Bemont
Radon	Rn	0	86	222.0	0.00973	?	The Curies, M. & Mme.
Rhenium	Re	—	75	—	?	?	Noddaek
Rhodium	Rh	3	45	102.91	12.5	1955.0	Wollaston
Rubidium	Rb	1	37	85.44	1.53	38.5	Bunsen, Kirchoff
Ruthenium	Ru	6, 8	44	101.7	12.2	2450	Claus
Samarium	Sa	3	62	150.43	7.7	1300	Boisbaudran
Scandium	Sc	3	21	45.10	2.5	1200	Wilson, Cleve
Selenium	Se	2, 4, 6	34	79.2	4.3	220	Berzelius
Silicon	Si	4	14	28.06	2.4	1420	Berzelius
Silver	Ag	1	47	107.88	10.5	960.5	—
Sodium	Na	1	11	23.0	0.97	97.5	Davy
Strontium	Sr	2	38	87.62	2.6	800.	Davy
Sulphur	S	2, 4, 6	16	32.06	1.96 to 2.07	106.8	—
Tantalum	Ta	5	73	181.5	16.6	2850	Eckeberg
Tellurium	Te	2, 4, 6	52	127.5	6.24	452.0	Von Reichenstein
Terbium	Tb	3	65	159.2	?	?	Mosander
Thallium	Tl	1, 3	81	204.4	11.85	302.0	Crookes
Thorium	Th	4	90	232.15	11.2	1845	Berzelius
Thulium	Tm	3	69	169.4	?	?	Cleve
Tin	Sn	2, 4	50	118.70	6.99	231.85	—
Titanium	Ti	4	22	47.9	4.5	1800.0	Gregor
Tungsten	W	6	74	184.0	19.3	3370.0	d'Elhujar
Uranium	U	4, 6	92	238.17	18.7	1850	Pelagot
Vanadium	V	3, 5	23	50.96	5.96	1710	Berzelius
Xenon	Xe	0	54	130.2	0.0058	- 140	Ramsay, Travers
Ytterbium	Yb	3	70	173.6	5.51	1800.0	Marignac
Yttrium	Yt	3	39	89.0	5.51	1490	Wöhler
Zinc	Zn	2	30	65.38	7.14	419.43	—
Zirconium	Zr	4	40	91.00	6.4	1700	Berzelius

SOME IMPORTANT MISCELLANEOUS ALLOYS

ALLOY OF	OTHER CONSTITUENTS	NAME OF ALLOY	PROPERTIES	USES
Aluminum	Mg, 5%-30%	Magnalium	Tough, strong, and light	Chem. balances; scientific instruments; airships
	Cu, 90%	Aluminum bronze	Golden; light; strong; inactive	Jewelry, statuary; scientific instruments; wire
	Cu, 5%	Duralumin	Silvery; light	Castings; household ware; aircraft
	Cu, 4%; Ni, 2%; Mg, 2%	Y alloy	Light; tough; inactive	Jewelry (looks like platinum when per cent of Ni is high)
Antimony	Pb, 82%; Sn, 5%	Type	Hard; expands on freezing	Printing
	Sn, 75%; Cu, 13%	Babbitt	Anti-friction; melts easily	Bearings
Bismuth	Pb, 25%; Sn, 12½%; Cd, 12½%	Wood's alloy	Melts at 60.5° C.	Automatic sprinklers and fire alarms; stereotype; fuses
	Pb, 40%; Sn, 20%	Rose's alloy	Melts at 94° C.	Boiler safety plugs; fuses
Copper	Zn, 30%-40%	Brass	Color; easy to cast; inactive	Doorknobs, etc.; plumbing fixtures, etc.
	Zn, 1%-25%; Sn, 1%-10%	Bronze	Inactive in air; casts well	Statuary, medals, etc.
	P, 1%; Pb, 10%; Sn, 10%	Phosphorbronze	Hard; acid resistant	Bearings; pumps; propellers
	Sn, 25%	Bell metal	Casts well	Bells
	Zn, 2.5%; Sn 2.5%	Cent		U. S. coin
	Zn, 20%; Ni, 30%	German silver (white metal)	Electric resistance; white color	Imitation silver; electric resistance
	Ni, 25%	Nickel coin	Color	U. S. coin
	Zn, 16%	Pinchbeck	Color of gold	Imitation gold jewelry
Chromium	Fe, 36%; Mo, 4%	Stainless steel	Inactive in acids	Joining two pieces of brass or copper
				Cutlery; vessels to hold corrosive liquids
Gold	Cu, 25%	18 carat gold	Harder than gold	Jewelry
	Cu, 10%	Gold coin	Harder than gold	Coins
	Ni, 20%-50%	White gold	Color; inactive; hard; like Pt.	Jewelry
	Ag, 70%	White gold	Color; inactive; hard; like Pt.	Jewelry
Iron	Ag, 12½%; Cu, 12½%; Cd, 12½%	Green gold	Color	Jewelry
	C, 5%; Si, 2%; (Mn, P, S, 0.7% each)	Cast iron	Cheap; casts well; hard	Stoves; large pipe; columns, railings, etc.
	C & Si, less than 1%	Wrought iron	Malleable; tough; welds; strength	Piano wire; eye-bars; sheet iron; nails, etc.
	C, ½-2%	Steel	Elastic; hard	Cutlery; I-beams, etc.; springs, etc.
	Cr, 1%-2%, C, 1%-2%	Chrome steel	Very hard; very strong	Armor plate; auto gears; rock crushers
	Mn, 11%	Manganese steel	Resists wear; tough	Shovels; crushers; rail curves; safes

SOME IMPORTANT MISCELLANEOUS ALLOYS—*Continued*

ALLOY OF	OTHER CONSTITUENTS	NAME OF ALLOY	PROPERTIES	USES
Iron	Mo, 3%-12%	Molybdenum steel	Retains temper at high heat	High speed lathe tools
	Ni, 2%-4%	Nickel steel	Strong; rusts less	Pumps; armor plate
	Si, 1%-5%	Silicon steel	Very magnetic	Dynamo and motor cores
	Si, 13%	Stainless steel	Inactive to acids, does not rust	Cutlery
	Si, 14%	Duriron	Hard; acid proof	Acid-resisting kettles and pipes
	V, 0.1-0.2%	Vanadium steel	Deoxidizes; strong; lasting	Axles, gears, railroad wheels, auto springs
	W, 6%-24%	Tungsten steel	Same as molybdenum steel	Same as molybdenum steel
Lead	Sn, 50%	Solder	Low melting pt., inactive	Joining metals
	Sb, 15%; Sn, 3%	Type	Hard, clear type	Printing
	As, $\frac{1}{2}$ %	Shot	Low melting pt., hard	Shot
	Alloys of Pb, Sn, Bi, and Cd	Fuses	Low melting pt.	Electric fuses; automatic fire sprinklers, alarms, etc.
Manganese	Fe, 80%-95%	Spiegel	Deoxidation; self-tempering	Bessemer and Open-Hearth steels
	Over 20% Mn, the rest is iron.	Ferromanganese	Deoxidation; self-tempering	Bessemer and Open-Hearth steels
Nickel	Cu, 75%	Nickel coin	Color and hardness	U. S. coin
	Cu, 28%; with little Fe & Mn	Monel metal	Acid resistance	Baskets for pickling metals; refrigerator linings
	Cr	Nichrome	High electric resistance and melting pt.	Electric toasters, flat irons, heaters
	Steel	Invar	Very small expansion	Clock pendulums; surveyors' instruments
	Fe, 58%	Platinite	Expands same as glass	Wires sealed in glass
Platinum	Ir, 5%	Platinum	Hard; high melting point	Jewelry
	Ag, 75%	Dental	Hard; inactive	Dentistry
	Ni, 18%; Cu, 18%; Cd, 1%	Watch metal	Hard; non-magnetic	Parts of watches
Silver	Cu, 10%	U. S. coin	Hard	U. S. coin; jewelry
	Cu, 7.5%	Sterling	Hard	Silverware; British coin
	Au, 30%	White gold	Color	Watch cases and other jewelry
Tin	Sb, 8%; Cu, 2%	Britannia (pewter)	White; hard, casts well	Table ware; ornamental fixtures
	Pb, 50%	Solder	Low melting point	Soldering; fusible metal

IMPORTANT TEMPERATURES

	DEGREES CENTIGRADE
Freezing point of helium	— 272
Boiling point of hydrogen	— 253
Boiling point of liquid air	— 192
Boiling point of oxygen	— 183
Freezing point of carbon dioxide	— 57
Melting point of ice	0
Melting point of Wood's alloy	60
Boiling point of water	100
Melting point of sulphur	107
Melting point of solder	190
Melting point of tin	232
Melting point of lead	327
Melting point of zinc	419
Boiling point of sulphur	445
Melting point of aluminum	660
Melting point of silver	960
Melting point of gold	1063
Melting point of copper	1083
Melting point of cast iron	1200
Melting point of steel	1250
Melting point of wrought iron	1600
Melting point of platinum	1755
Hottest part of the Bunsen burner	1870
Oxy-hydrogen flame	2800
Oxy-acetylene flame	3500
Electric arc furnace	3500
Sun's temperature	more than 5000

NOBEL PRIZES IN CHEMISTRY

YEAR	WINNER	ACHIEVEMENT
1901	J. H. van't Hoff	Nature of solution
1902	Emil Fischer	Chemistry of the sugars
1903	S. Arrhenius	Theory of electrolytic dissociation
1904	William Ramsay	Rare gases of the air
1905	A. von Baeyer	Synthesis of indigo
1906	H. Moissan	Isolation of fluorine
1907	E. Buchner	Nature of zymase (enzyme) in yeast
1908	E. Rutherford	Radio-activity
1909	Wm. Ostwald	Fixation of nitrogen (HNO_3)
1910	O. Wallach	Chemistry of the terpenes
1911	Marie Curie	Radium
1912	V. Grignard	Organic synthesis
	P. Sabatier	Organic synthesis
1913	A. Werner	Structure of complex compounds
1914	T. W. Richards	Atomic weight determinations
1915	R. Willstätter	Chemistry of chlorophyll
1916	Not awarded	
1917	Not awarded	
1918	F. Haber	Fixation of nitrogen (NH_3)
1919	Not awarded	
1920	W. Nernst	Physical chemistry
1921	F. Soddy	Radio-active elements
1922	F. W. Aston	Isotopes
1923	F. Pregl	Micro-analysis
1924	Not awarded	
1925	R. Zsigmondy	Ultra-microscope
1926	T. Svedberg	Colloids
1927	H. Wieland	Bile acids
1928	A. Windaus	Vitamins

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APPROXIMATE ATOMIC WEIGHTS OF THE COMMON ELEMENTS

NAME	SYMBOL	APPROXIMATE ATOMIC WT.
Aluminum	Al	27
Barium	Ba	137
Bromine	Br	80
Calcium	Ca	40
Carbon	C	12
Chlorine	Cl	35.5
Copper	Cu	64
Fluorine	F	19
Gold	Au	197
Hydrogen	H	1
Iodine	I	127
Iron	Fe	56
Lead	Pb	207
Magnesium	Mg	24
Manganese	Mn	55
Mercury	Hg	200
Nitrogen	N	14
Oxygen	O	16
Phosphorus	P	31
Potassium	K	39
Silicon	Si	28
Silver	Ag	108
Sodium	Na	23
Sulphur	S	32
Tin	Sn	119
Zinc	Zn	65

